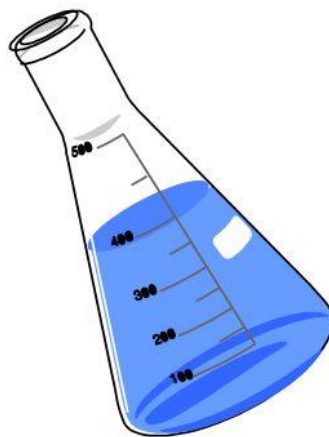




ACIDS AND BASES



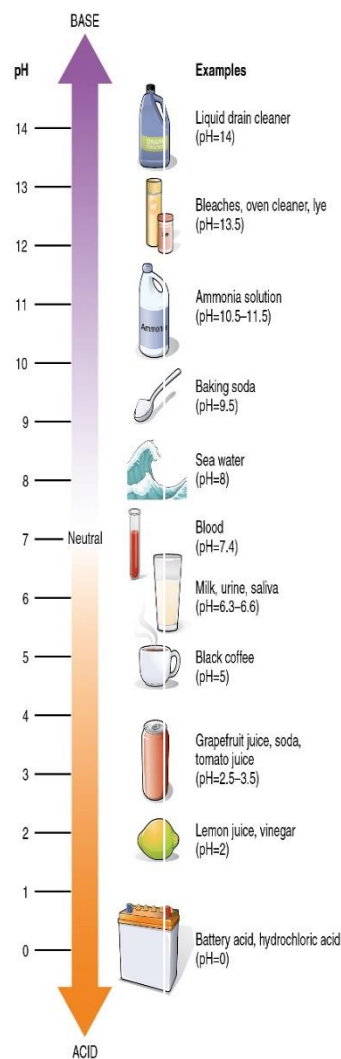
Grade 12 advanced

2024/2025

Prepared by : Naeem El Emam Akl

Student name:.....

Class :.....



هذه المذكرة لاتعني عن الكتاب المدرسي وليست بهدف البيع او الربح

تمنيتي لجميع الطلبة والطالبات بالتفوق والنجاح

item	Definetions
Achdic solutions	Asolution contains more hydrogen ions than hydroxide ions
Abasic solution	Asolution contains more hydroxide ions than hydrogen ions
Conjugate acids	Is the species produced when a base accepted ahydrogen ion
Self ionization of water	I which water molecules react with other water molecules as Bronsted- Lowry model to form a hydronium ion H_3O^+ and a hydroxide ion OH^- $\text{H}_2\text{O}_{(\text{L})} + \text{H}_2\text{O}_{(\text{L})} \rightleftharpoons \text{H}_3\text{O}^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$ Or $\text{H}_2\text{O}_{(\text{L})} \rightleftharpoons \text{H}^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$
Conjugate base	Is the species that results when an acid donates ahydrogen ion
Aconjugate acid-base pair	Consists of two substances related to each other by the donating and accepting of asingle hydrogen ion
Amphoteric compound	That compound can act as both acids and bases
Mono protic acid	An acid that can donate only one hydrogen ion per molecule
Diprotic acids	Acids that contain two ionizable hydrogen atoms per molecule
Tri protic acids	Acids that contain three ionizable hydrogen atoms per molecule
Poly protic acids	Acids that have more than one ionizable hydrogen atoms
stalactites	Deposits shaped like icicles that form on the ceiling
stalagmites	Rounded masses rising from the floor

Notes

1-you should never attempt to identify an acid or abase or any other substance in the laboratory by its taste or feel

2-rhododendrons flourish in rich moist soil that is moderately acidic , whereas sempervivum commonly called hen and chicks grows in best in drier slightly basic soil

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Properties of acids and bases:

acids	Bases
✓ Taste sour ✓ React with litmus and cause blue litmus paper to turn red ✓ React with metals ✓ Acid + active metal $\rightarrow$ H₂ gas $\text{Zn}_{(s)} + 2\text{HCl}_{(aq)} \rightarrow \text{ZnCl}_{2(aq)} + \text{H}_{2(g)}$ $2\text{Al}_{(s)} + 3\text{H}_2\text{SO}_{4(aq)} \rightarrow \text{Al}_2(\text{SO}_4)_3(aq) + 3\text{H}_{2(g)}$ ✓ React with metal carbonate AND Hydrogen carbonates: ✓ Acid + metal carbonate $\rightarrow$ salt + H₂O + CO₂ $\text{NaHCO}_{3(s)} + \text{HC}_2\text{H}_3\text{O}_{2(aq)} \rightarrow \text{NaC}_2\text{H}_3\text{O}_{2(aq)} + \text{H}_2\text{O}_{(l)} + \text{CO}_{2(g)}$ ✓ PH < 7	✓ Taste bitter and feel slippery ✓ React with litmus and cause red litmus paper to turn blue اعداد أ. نعيم الامام عقل 0508194296 ✓ PH > 7
✓ Acid and base solution are conduct electricity(why?) Because they have free moving ions that cause the resulting solution to become a conductor ✓ They change color of indicators	

acids	Uses
Formic acid (methanoic acid) HCOOH	✓ Ants emit it when they sense danger to alert their entire ant colony ✓ used to process latex tapped from rubber trees into natural rubber
Carbonic acid H₂CO₃ Phosphoric acid H₃PO₄	✓ Give many carbonated beverages their sharp taste
Citric acid Ascorbic acid (vitamin c)	✓ Give lemons and grape fruit their tartness
Acetic acid (ethanoic acid) CH₃COOH	✓ Makes vinegar taste sour
Hydrochloric acid HCl (muriatic acid)	✓ It is found in stomach to help digest food ✓ Clean bricks and concrete ✓ Identify rocks as limestone (primarily CaCO₃) if a few drops of acid produce bubbles of carbon dioxide (CO₂) the rocks contain limestone
Hydrocyanic acid (prussic acid) HCN	✓ Dyeing, Engraving and Tempering steel ✓ Hydrogen cyanide is a poisonous gas found in exhaust of vehicles, in tobacco, wood smoke and in smoke from burning nitrogen-containing plastics. ✓ Some insects such as millipedes and burnet moths release it as a gas to defense mechanism ✓ Fruits that have a pit such as cherries or peaches contain cyanohydrins which convert to hydrocyanic acid in the digestive system if the pit is eaten- however no hydrocyanic acid is produced in the flesh of these fruits so the fruit can safely be eaten

Bases	Uses
Ammonia solutions (NH_3)	✓ Household products (cleaner) ✓ Farmers are able to increase the nutritional value of the low quality vegetable materials such as straw . hay and other crops residue by immersing the materials in an atmosphere of ammonia gas for 3 weeks
Magnesia milk ($\text{Mg}(\text{OH})_2$)	✓ Anti acids tablet
Sodium hydroxide (NaOH)	✓ Clear clogged drains ✓ Make soap

The Arrhenius model

Arrhenius acid	Arrhenius base
✓ Is a substance that contains hydrogen and ionizes to produce (H^+) ions in aqueous solution which make the solution acidic ✓ Example: $\text{HCl}_{(g)} \rightarrow \text{H}^+_{(aq)} + \text{Cl}^-_{(aq)}$	✓ Is a substance that contains a hydroxide group (OH^-) and dissociates to produce hydroxide ion in aqueous solution ✓ Example: $\text{NaOH}_{(s)} \rightarrow \text{Na}^+_{(aq)} + \text{OH}^-_{(aq)}$

Arrhenius Succeeded in :

Explaining many acidic and basic solutions

Fall in (shortcoming) :

Some compounds don't contain hydroxide group yet both substances produce hydroxide ions in solution and are well known bases like:

- ✓ Ammonia NH_3
- ✓ Na_2CO_3

Notes

- ✓ Na_2CO_3 is the compound that causes the alkalinity of lake natron Tanzania
- ✓ Lake natron in Africa's great rift valley is a naturally basic body of water. Water, laden with dissolved sodium carbonate from surrounding volcanic rocks, drains into the lake but finds no outlet. evaporation concentrates the mineral leaving a white crust on the surface and strongly alkaline water

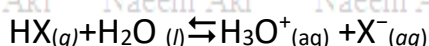
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The Bronsted- lowry model

- ✓ This model more inclusive model of acids and bases
- ✓ This model focuses on the hydrogen ion (H⁺)
- ✓ Acid : is a hydrogen ion donor
- ✓ A base : is a hydrogen ion acceptor

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Hydrogen ion doners and acceptors :

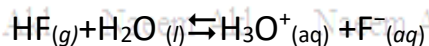
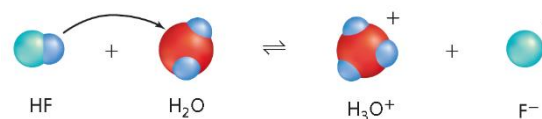


acid base Conjugate Acid Conjugate base

The symbols X,Y represent: non metallic element or negative poly atomic ions

Conjugate acids and base

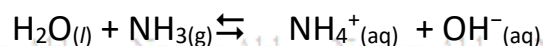
➤ Hydrogen fluoride – Bronsted-Lowry acid :



acid base Conjugate Acid Conjugate base

- ✓ Hydrogen fluoride react with organic compounds called hydrocarbons to substitute fluorine atoms for hydrogen atoms

➤ Ammonia – Bronsted-Lowry base :



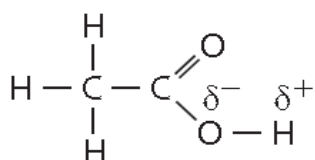
acid base Conjugate Acid Conjugate base

➤ Water- a Bronsted- Lowry acid and base (amphoteric compound) :

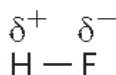
Water can act as either an acid or a base depending on what other substances are in solution

➤ **A Hydrogen is ionizable depends on the polarity of its bond**

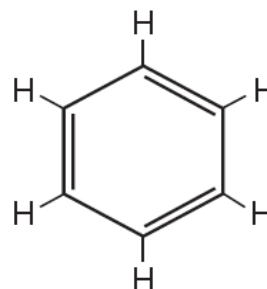
Compound	Bronsted- lowry acid	Illustrate
Acetic acid CH_3COOH	YES	✓ Because oxygen is more electronegative than hydrogen ✓ The bond between hydrogen and oxygen is polar ✓ So the hydrogen atom can ionize in solution
Hydrogen fluoride in solution HF	YES	➤ Fluorine is highly electronegative so HF is an acid in solution
Benzene C_6H_6	NO	➤ Because the difference between carbon and hydrogen atoms little electronegativity



Acetic acid



Hydrogen fluoride



Benzene

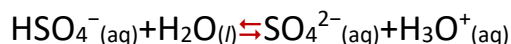
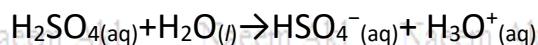
➤ **Monoprotic and polyprotic acids:**

Monoprotic acids	polyprotic acids	
	Diprotic acids	Triprotic acids
✓ An acid that can donate only one hydrogen ion per molecule (H^+) ✓ Examples: Phosphoric acid H_3PO_4 Boric acid H_3BO_3	✓ Acids that contain two ionizable hydrogen atoms per molecule (2H^+) ✓ Examples: Sulphuric acid H_2SO_4 Carbonic acid H_2CO_3 Hydrosulphuric acid H_2S	✓ Acids that contain three ionizable hydrogen atoms per molecule (3H^+) ✓ Examples: Perchloric acid HClO_4 Hydrochloric acid HCl Nitric acid HNO_3 Acetic acid CH_3COOH

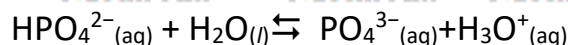
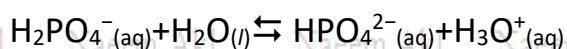
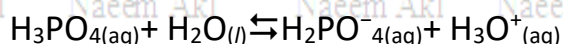
Acid		Conjugate Base	
Name	Formula	Name	Formula
Hydrochloric acid	HCl	Chloride ion	Cl ⁻
Nitric acid	HNO ₃	Nitrate ion	NO ₃ ⁻
Sulfuric acid	H ₂ SO ₄	Hydrogen sulfate ion	HSO ₄ ⁻
Hydrogen sulfate ion	HSO ₄ ⁻	Sulfate ion	SO ₄ ²⁻
Hydrofluoric acid	HF	Fluoride ion	F ⁻
Hydrocyanic acid	HCN	Cyanide	CN ⁻
Acetic acid	HC ₂ H ₃ O ₂	Acetate ion	C ₂ H ₃ O ₂ ⁻
Phosphoric acid	H ₃ PO ₄	Dihydrogen phosphate ion	H ₂ PO ₄ ⁻
Dihydrogen phosphate ion	H ₂ PO ₄ ⁻	Hydrogen phosphate ion	HPO ₄ ²⁻
Hydrogen phosphate ion	HPO ₄ ²⁻	Phosphate ion	PO ₄ ³⁻
Carbonic acid	H ₂ CO ₃	Hydrogen carbonate ion	HCO ₃ ⁻
Hydrogen carbonate ion	HCO ₃ ⁻	Carbonate ion	CO ₃ ²⁻

✓ Diprotic acids ionize in steps :

The two ionizations of sulphuric acid are described by the following equations:



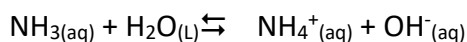
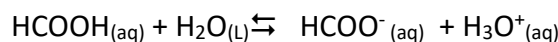
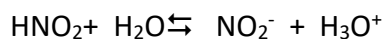
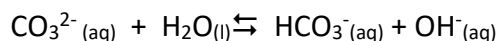
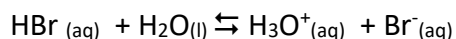
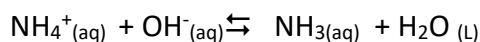
The three ionizations phosphoric acid are described by the following equations:



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1-Identify the conjugate acid – base pairs in each reaction



2-challenge: the products of an acid – base reaction are H₃O⁺ and SO₄²⁻

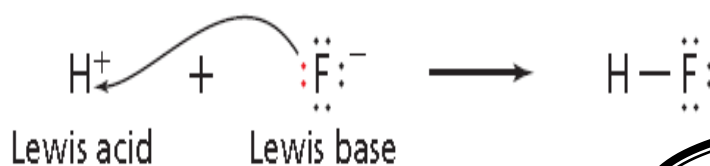
Write a balanced equation for the reaction and identify the conjugate acid – base

The Lewis model

- ✓ Lewis developed the electron-pair theory of the chemical bonding and introduced lewis structures to keep track of the electrons in the atoms and molecules
- ✓ He applied his electron-pair theory of the chemical bonding to acid-base reactions
- ✓ **Acid:** is an ion or molecule with a vacant atomic orbital that can accept (share) an electron pair [Lewis acid: is an electron pair acceptor]
- ✓ **Base:** is an ion or molecule with with a lone electron pair that it can donate(share) [Lewis base: is an electron pair doner]
- ✓ The lewis model includes all substances classified as bronsted- lowry acids and bases and many more.

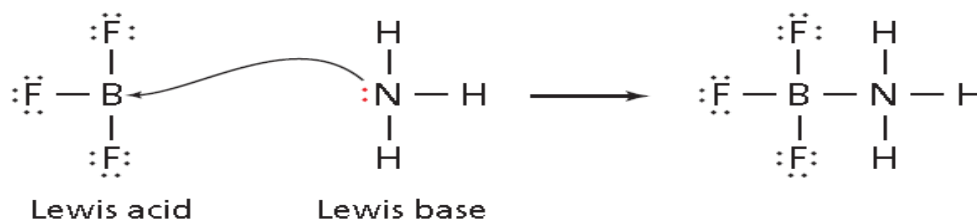
➤ Electron pair doner and acceptors:

- ✓ **Formation of hydrogen fluoride :**



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- ✓ The hydrogen ion (H^+) is the Lewis acid
- ✓ Its vacant 1s orbital accepts an electron pair from the F^- ion
- ✓ The F^- is the Lewis base it donates alone electron pair to form hydrogen-fluorine bond in HF
- ✓ This reaction conforms to the Bronsted- Lowry model of acids and bases because (H^+) can be considered a hydrogen ion donner and F^- a hydrogen ion acceptor
- ✓ **reaction between BF_3 and NH_3 to form BF_3NH_3 :**
- ✓ The boron atom (B) in BF_3 has 6 valence electrons
- ✓ Vacant orbital can accept an electron pair from a lewis base
- ✓ **Reaction between gaseous sulfur trioxide (SO_3) with magnesium oxide (MgO)**



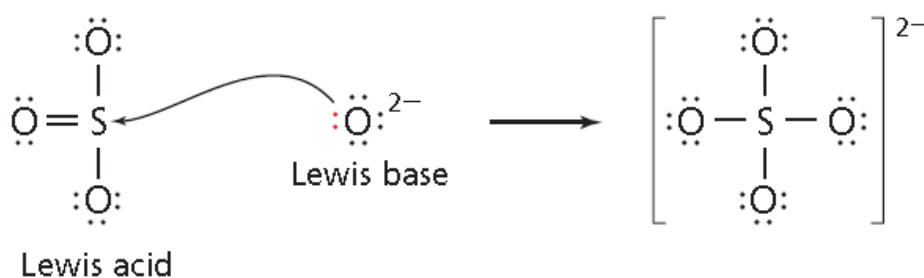


- ✓ The acid base part of the reaction involves SO_3 and oxide ion (O^{2-}) of MgO
- ✓ SO_3 is a Lewis acid accepts an electron pair from O^{2-} ion a Lewis base

The reaction between SO_3 with MgO is important

- 1- because it produces magnesium sulphate salt that forms **Epsom salt**
- 2- **has environmental application.**

- ✓ When MgO injected into the flue gases of coal-fired power plants
- ✓ It reacts with and removes SO_3 (a waste gas from the burning of coal can



removed from smokestack gases) because if SO_3 is allowed to enter the atmosphere it can combine with water in the air to form sulfuric acid which falls to earth as acid precipitation.

- ✓ Note that while there is a good deal of steam coming from the cooling towers there is little visible smoke from the smokestack.

Epsom salt ($\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$):

Magnesium sulphate heptahydrate

Uses:

- ✓ including soothing sore muscles
- ✓ Acting as a plant nutrient

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model	Acid definition	Base definition
arrhenius	H^+ Producer	OH^- Producer
Bronsted - lowry	H^+ donor	H^+ acceptor
lewis	Electron-pair acceptor	Electron-pair doner

✓ Anhydrides:

Acid anhydride	An alkaline anhydride
✓ is an oxide that can combine with water to form an acid ✓ Oxides of nonmetals form acids ✓ Make acidic rain water ✓ Examples: $\text{CO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{CO}_3$ $\text{SO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_4$	✓ an oxide that can combine with water to form the bases ✓ Oxides of metallic elements form bases ✓ Examples: (lime) $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$

✓ Acidic rain water :

Causes: reaction between non-metallic oxide with water to form acid

Results:

- ✓ When it reaches the ground some sinks into the soil and reaches limestone bedrock where dissolves the lime stone
- ✓ The dissolution of lime stone creates huge underground caverns
- ✓ Within a cavern, groundwater might drip from the ceiling and deposit some of dissolved limestone (**stalactites- stalagmites**)

- 1- Write the lewis structure for phosphorus tri chloride (PCl_3) , is PCl_3 a lewis acid , lewis base, neither
- 2- In the accompanying structural formula , identify any hydrogen atoms that are likely to be ionizable

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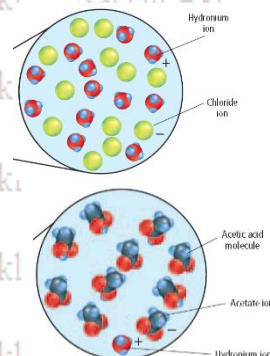
Section 2

Strengths of acids and bases

Strong acids	Weak acid
✓ Acids that ionize completely ✓ Solutions have ions only ✓ Strong electrolyte (its aqueous solutions conduct electricity) ✓ Example: $\text{HCl}_{(g)} + \text{H}_2\text{O}_{(l)} \rightarrow \text{H}_3\text{O}^+_{(aq)} + \text{Cl}^-_{(aq)}$	✓ Acids that ionizes only partially in dilute aqueous solutions ✓ Solutions have molecules and ions ✓ Weak electrolyte ✓ Example: $\text{CH}_3\text{COOH}_{(aq)} + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{H}_3\text{O}^+_{(aq)} + \text{CH}_3\text{COO}^-_{(aq)}$

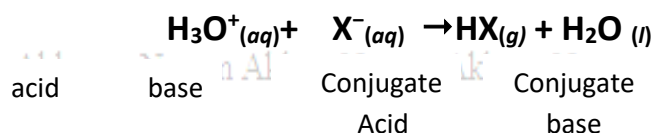
Note

- ✓ The glow of the bulb indicates that the solution conduct electricity
- ✓ The brightness of the bulb connected to the HCl is high
- ✓ The **light glows brightly** when electrodes are placed in 0.1M HCl (why?)
Because all of the HCl is in the form (H_3O^+ , Cl^-)
- ✓ The **light is dim** when electrodes are placed in 0.1M $\text{HC}_2\text{H}_3\text{O}_2$ (why?)
Because acetic acid produce fewer ions ($\text{HC}_2\text{H}_3\text{O}_2$, H_3O^+ , acetate ion)



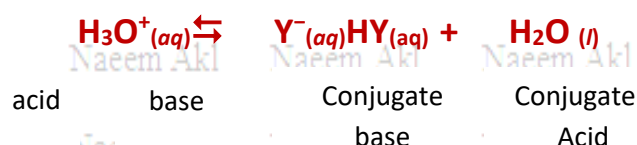
Strong Acids		Weak Acids	
Name	Ionization Equation	Name	Ionization Equations
Hydrochloric	$\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$	Hydrofluoric	$\text{HF} \rightleftharpoons \text{H}^+ + \text{F}^-$
Hydroiodic	$\text{HI} \rightarrow \text{H}^+ + \text{I}^-$	Acetic	$\text{HC}_2\text{H}_3\text{O}_2 \rightleftharpoons \text{H}^+ + \text{C}_2\text{H}_3\text{O}_2^-$
Perchloric	$\text{HClO}_4 \rightarrow \text{H}^+ + \text{ClO}_4^-$	Hydrosulfuric	$\text{H}_2\text{S} \rightleftharpoons \text{H}^+ + \text{HS}^-$
Nitric	$\text{HNO}_3 \rightarrow \text{H}^+ + \text{NO}_3^-$	Carbonic	$\text{H}_2\text{CO}_3 \rightleftharpoons \text{H}^+ + \text{HCO}_3^-$
Sulfuric	$\text{H}_2\text{SO}_4 \rightarrow \text{H}^+ + \text{HSO}_4^-$	Hypochlorous	$\text{HClO} \rightleftharpoons \text{H}^+ + \text{ClO}^-$

✓ Acid strength and the Bronsted-lowry model :



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- ✓ HX :represents a strong acid (is nearly 100% ionized) and its conjugate base is weak
- ✓ The ionization equilibrium lies almost completely to the right (why?)
Because the base H_2O (**stronger base in the forward reaction**) has much greater attraction for the H^+ ion than does the conjugate base X^- (**in the reverse reaction**)



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✓ **HY** represents a weak acid

✓ The ionization equilibrium for a weak acid lies far to the left (why?)

The conjugate base Y^- has greater attraction for the H^+ ion than does the base H_2O ✓ In the battle of the bases the conjugate base Y^- is stronger than the base H_2O and manages to capture the H^+ ion✓ **Strengths of bases :**

Strong base	Weak base
✓ A base that dissociates entirely into metal ions and hydroxide ions (OH^-) ✓ Examples: $\text{NaOH}_{(\text{s})} \rightarrow \text{Na}^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$ ✓ Some metallic hydroxides such as $\text{Ca}(\text{OH})_2$ have low solubility (solubility product constant K_{sp} is small 6.5×10^{-6} <i>indicating the few OH^- ions are present in a saturated solution</i>) and thus are poor sources of OH^- ions $\text{Ca}(\text{OH})_2 \rightleftharpoons \text{Ca}^{2+}_{(\text{aq})} + 2\text{OH}^-_{(\text{aq})}$ ✓ Strong electrolyte	✓ Bases that ionize only partially in dilute aqueous solutions ✓ Examples: $\text{CH}_3\text{NH}_2_{(\text{g})} + \text{H}_2\text{O}_{(\text{l})} \rightleftharpoons \text{CH}_3\text{NH}_3^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$ base acid acid Conjugate base ✓ The equilibrium lies far to the left because the base CH_3NH_2 is weak and the conjugate base OH^- is strong – the OH^- has a greater attraction for a hydrogen ion than has a molecule of methyl amine ✓ Weak electrolyte

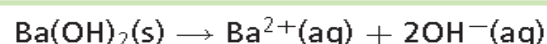
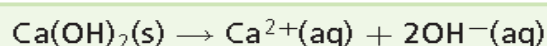
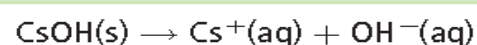
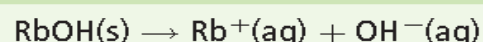
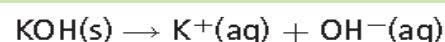
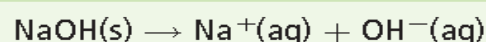
Not

✓ Nevertheless, $\text{Ca}(\text{OH})_2$ and other slightly soluble metallic hydroxides are considered strong bases (why?)

Because all of the compound that dissolves is completely dissociation

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**Dissociation Equations for Strong Bases**

✓ Acid ionization constants and Base ionization constants

Acid ionization constants (K_a)	Base ionization constants (K_b)
✓ Is the value of the equilibrium constant expression for the ionization of a weak acid ✓ Example: $\text{HCN}_{(g)} + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{H}_3\text{O}^+_{(aq)} + \text{CN}^-_{(aq)}$ $K_{eq} = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}][\text{H}_2\text{O}]}$ $K_a = K_{eq}[\text{H}_2\text{O}]$ $K_a = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}]} = 6.2 \times 10^{-10}$	✓ Is the value of the equilibrium constant expression for the ionization of a weak base ✓ Example: $\text{CH}_3\text{NH}_2_{(g)} + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{CH}_3\text{NH}_3^+_{(aq)} + \text{OH}^-_{(aq)}$ $K_b = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]}$ ✓ The smaller the value of K_b the weaker the base

Note

- ✓ although the Bronsted – lowry model helps explain acid strength, the model does not provide a quantitative way to express the strength of acid or to compare the strengths of various acids
- ✓ The equilibrium constant expression provides the quantitative measure of acid strength (measure the degree of ionization of the acid)
- ✓ The value of K_a indicates whether reactants or products are favored at equilibrium
- ✓ **The weakest acids** have the **smallest K_a** values because their solutions have the **lowest** concentration of **ions** (products in the numerator) and the **highest** concentration of **un-ionized acid molecules** (reactant in the denominator)
- ✓ Polyprotic acids are not necessarily strong acids for any of their ionizations, each ionization of a polyprotic acid has a K_a value and the value decreases for each successive ionization

Ionization Constants for Weak Acids		
Acid	Ionization Equation	K_a (298 K)
Hydrosulfuric, first ionization	$\text{H}_2\text{S} \rightleftharpoons \text{H}^+ + \text{HS}^-$	8.9×10^{-8}
Hydrosulfuric, second ionization	$\text{HS}^- \rightleftharpoons \text{H}^+ + \text{S}^{2-}$	1×10^{-19}
Hydrofluoric	$\text{HF} \rightleftharpoons \text{H}^+ + \text{F}^-$	6.3×10^{-4}
Hydrocyanic	$\text{HCN} \rightleftharpoons \text{H}^+ + \text{CN}^-$	6.2×10^{-10}
Acetic	$\text{CH}_3\text{COOH} \rightleftharpoons \text{H}^+ + \text{CH}_3\text{COO}^-$	1.8×10^{-5}
Carbonic, first ionization	$\text{H}_2\text{CO}_3 \rightleftharpoons \text{H}^+ + \text{HCO}_3^-$	4.5×10^{-7}
Carbonic, second ionization	$\text{HCO}_3^- \rightleftharpoons \text{H}^+ + \text{CO}_3^{2-}$	4.7×10^{-11}

Ionization Constants of Weak Bases

Base	Ionization Equation	K_b (298 K)
Ethylamine	$C_2H_5NH_2(aq) + H_2O(l) \rightleftharpoons C_2H_5NH_3^+(aq) + OH^-(aq)$	5.0×10^{-4}
Methylamine	$CH_3NH_2(aq) + H_2O(l) \rightleftharpoons CH_3NH_3^+(aq) + OH^-(aq)$	4.3×10^{-4}
Ammonia	$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$	2.5×10^{-5}
Aniline	$C_6H_5NH_2(aq) + H_2O(l) \rightleftharpoons C_6H_5NH_3^+(aq) + OH^-(aq)$	4.3×10^{-10}

PRACTICE PROBLEMS

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3- write the ionization equations and acid ionization constant expressions for each acid :

- ✓ $HClO_2$
- ✓ HNO_2
- ✓ HIO

4- write the first and second ionization equations for H_2SeO_3

5-Challenge : given the expression $K_a = \frac{[ASO_4][H_3O]}{[HASO_4]}$, write the balanced equation for the corresponding reaction ?

6- write the ionization equations and base ionization constant expressions for each base

- a- Hexyl amine ($C_6H_{13}NH_2$)
- b- Propyl amine ($C_3H_7NH_2$)
- c- Carbonate ion (CO_3^{2-})
- d- Hydrogen sulphite ion (HSO_3^-)
- e- Aniline ($C_6H_5NH_2$)

7-challenge: write an equation for a base equilibrium in which the base in the forward reaction is PO_4^{3-} and the base in the reverse reaction is OH^-

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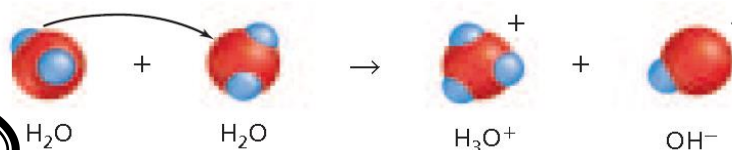
Section 3

Hydrogen ion and PH

MAIN IDEA:

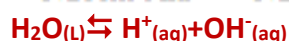
PH and POH are logarithmic scales that express the concentration of hydrogen ion and hydroxide ion in aqueous solutions

Item	Definition
The ion product constant for water	Is the value of the equilibrium constant expression for the self-ionization of water and equal constant value at 25°C (298K) 1×10^{-14}
The PH of solution	Is the negative logarithm of the hydrogen ion concentration $\text{pH} = -\log[\text{H}_3\text{O}^+]$
The POH of solution	Is the negative logarithm of the hydroxide ion concentration $\text{pOH} = -\log[\text{OH}^-]$
Molarity	Is the number of moles of molecules or formula units that were dissolved in 1L of solution
A strip of PH paper	Is dipped into an acidic solution or a basic solution the color of the paper changes (PH papers treated with several indicators can identify a range of PH value)
The PH meter	That meter provides a more accurate measure of PH when electrodes are placed in a solution and gives a direct analog or digital readout of PH

✓ **Ion product constant for water (K_w) :**

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- ✓ **Pure Water** is neutral because that contains equal concentration of H^+ and OH^- ions produced by self ionization (number of hydronium equal number of hydroxide ions)



- ✓ in pure water at 298 K: $[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{M}$
- ✓ the concentration of pure water is constant so the $[\text{H}_2\text{O}]$ does not appear in the denominator of K_w expression
- ✓ K_w is a special equilibrium constant expression that applies only to water
- ✓ **$K_w = [\text{OH}^-] \times [\text{H}^+] = (1.0 \times 10^{-7} \text{M}) \times (1.0 \times 10^{-7} \text{M}) = 1.0 \times 10^{-14}$ at 298K**

The Ion Product of Water

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

K_w is the ion product constant for water. $[\text{H}^+]$ represents the concentration of the hydrogen ion. $[\text{OH}^-]$ represents the concentration of the hydroxide ion.

In dilute aqueous solutions, the product of the concentrations of the hydrogen ion and the hydroxide ion equals K_w .

✓ Kw and Le Chatelier's principle :

If the concentration of H^+ ions increase the concentration of OH^- must decrease to keep the value of K_w constant (1.0×10^{-14} at 298K)

If adding extra hydrogen ions to water at equilibrium is a stress on the system

- the system reacts in a way to relieve the stress
- the added H^+ ions react with OH^- ions to form more water molecules
- thus the concentration of OH^- ions decreases

example 1:

calculate $[H^+]$, $[OH^-]$ at 298K the H^+ ion concentration in a cup of coffee is $1 \times 10^{-5} M$ what is the OH^- ion concentration in the coffee? is the coffee acidic or basic or neutral?

Solve :

$[H_3O^+] =$ concentration of strong acid

$$[H_3O^+] = 1.0 \times 10^{-5} M$$

$$K_w = [OH^-] \times [H_3O^+] = 1.0 \times 10^{-14}$$

$$[OH^-] = \frac{1.0 \times 10^{-14}}{[H^+]} = \frac{1 \times 10^{-14}}{1.0 \times 10^{-5}} = 1 \times 10^{-9} mol/L$$

$[OH^-]$ because $<$ The Coffee is acidic $[H_3O^+]$

PRACTICE Problems:

8- complete the following table :

$[H^+]$	$[OH^-]$	نوع المحلول	PH
$1.0 \times 10^{-13} M$			
	$1.0 \times 10^{-3} M$		
$1.0 \times 10^{-7} M$			
	$4.0 \times 10^{-5} M$		
$1.0 \times 10^{-2} M$			
$3.0 \times 10^{-6} M$			
0.0055M			
0.000084M			

9- challenge calculate the number of H^+ ions and the number of OH^- ions in 300 ml of pure water at 298 K?

10- calculate $[H^+]$, $[OH^-]$ at 298K in a sample of seawater with a $POH = 5.60$?

At 298K

$$K_w = [OH^-] \times [H^+] = 1 \times 10^{-14}$$

$$pH = -\log[H_3O^+]$$

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

$$pH + pOH = 14$$

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PH and POH at 298K

Acidic solution	Basic solution	Neutral solution
$[H_3O^+] > [OH^-]$ $[OH^-] < [H_3O^+]$ $PH < 7$ $POH > 7$ $POH > PH$	$[H_3O^+] < [OH^-]$ $[OH^-] > 1 \times 10^{-7}$ $PH > 7$ $POH < 7$ $PH > POH$	$[OH^-] = [H_3O^+] = 1 \times 10^{-7} M$ $pH = 7$ $pOH = 7$

- ✓ A solution with a PH of 0.0 is strongly acidic
- ✓ A solution with a PH of 14.0 is strongly basic
- ✓ A solution having a PH of 3.0 has ten times the hydrogen ion concentration of a solution with a PH of 4.0
- ✓ The logarithmic nature of the PH scale means that a change of one PH unit represents a tenfold change in ion concentration

pH

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

$[H^+]$ represents the hydrogen ion concentration.

The pH of a solution equals the negative logarithm of the hydrogen ion concentration.

pOH

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

$[OH^-]$ represents the hydroxide ion concentration.

The pOH of a solution equals the negative logarithm of the hydroxide ion concentration.

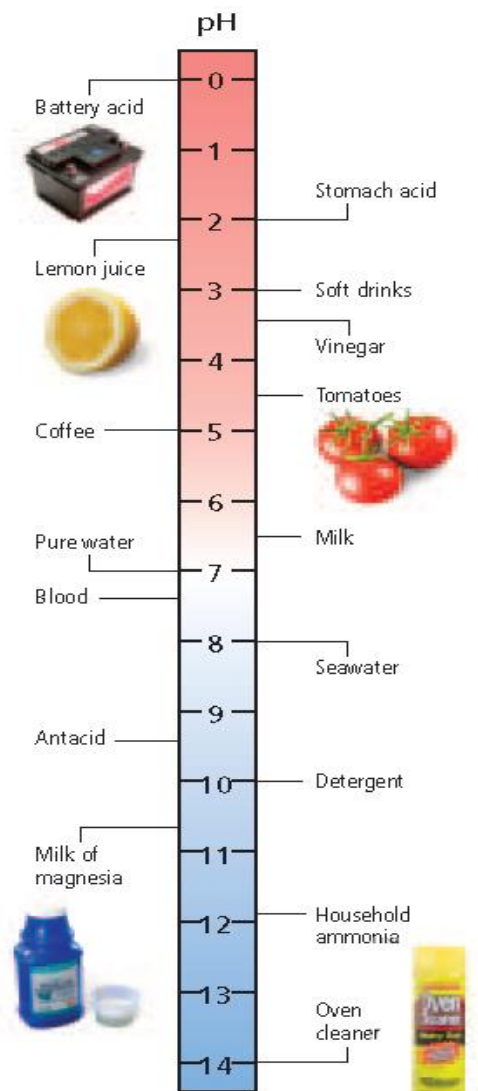
How pH and pOH Are Related

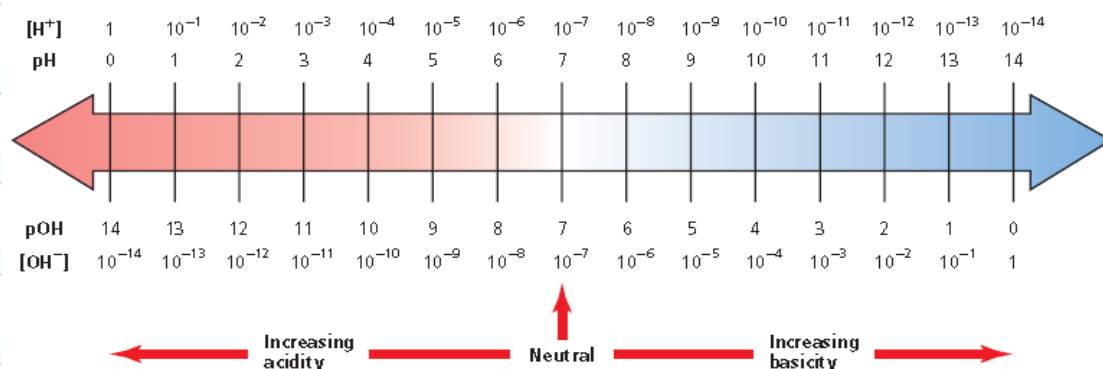
$$pH + pOH = 14.00$$

pH represents $-\log [H^+]$.
pOH represents $-\log [OH^-]$.

The sum of pH and pOH is 14.00.

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example 2:

cleaner has hydroxide ion of 4×10^{-3} M

calculate the POH , PH of cleaner at 298K ?

SOLVE:

$$\therefore [\text{OH}^-] = 4 \times 10^{-3} \text{ M}$$

$$K_w = [\text{OH}^-] \times [\text{H}_3\text{O}^+] = 1.0 \times 10^{-14}$$

$$[\text{H}_3\text{O}^+] = K_w \div [\text{OH}^-] = 1.0 \times 10^{-14} \div 4.0 \times 10^{-3} = 0.25 \times 10^{-11} \text{ M}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log 4 \times 10^{-3} = 2.4$$

$$\text{PH} = 14 - 2.4 = 11.60$$

example 3:

What are $[\text{H}^+]$, $[\text{OH}^-]$ at 298K in a healthy persons blood that has a PH of 7.40 ?

solve:

$$[\text{H}_3\text{O}^+] = 10^{-\text{PH}} = 4.0 \times 10^{-8} \text{ M}$$

$$[\text{H}_3\text{O}^+] = 10^{-7.4}$$

$$\text{PH} + \text{POH} = 14$$

$$\text{POH} = 14 - 7.4 = 6.60$$

$$[\text{OH}^-] = 10^{-6.6} = 2.5 \times 10^{-7} \text{ M}$$

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✓ Molarity and the PH of strong acids :

- ✓ for all strong monoprotic acids the concentration of the acid is the concentration of H^+ ion so we can use the molarity of the acid to calculate PH

✓ example:



✓ What's the meaning 0.1M HCl ?

That mean 0.1 mol of H^+ ion per liter and 0.1 mol of Cl^- ions per liter because HCl is essentially 100% in the form ions in solution

✓ Molarity and the PH of strong bases :

- ✓ Strong bases are fully ionized
- ✓ One formula unit of NaOH produces OH⁻ ion thus the concentration of the OH⁻ ions is the same as molarity of the solution 0.1M NaOH



- ✓ The concentration of OH⁻ ion in each formula unit 0.1M Ca(OH)₂ is twice of the molarity of ionic compound

$$[\text{OH}^-] = 2 \times 0.1\text{M} = 0.2\text{M OH}^-$$

- ✓ You must use K_a, K_b values to determine the concentration of H⁺ and OH⁻ ions in solution of Weak acids and bases because Weak acids and weak bases are only partially ionized

Calculating Ka from PH :

example 4:

If the concentration HF=0.100M and PH=3.20 calculate Ka of HF ?

Solve:



$$K_a = \frac{[\text{H}][\text{F}]}{[\text{HF}]}$$

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$$[\text{H}^+] = 10^{-\text{PH}} = 10^{-3.2} = 6.13 \times 10^{-4} \text{ M}$$

at equilibrium : $[\text{HF}] = 0.1 - [\text{H}^+] = 0.1 - 6.13 \times 10^{-4} = 0.0994 \text{ M}$

$$K_a = \frac{[\text{H}][\text{F}]}{[\text{HF}]} = \frac{(6.13 \times 10^{-4})^2}{0.0994} = 3.793 \times 10^{-6}$$

Practice problems :

11-calculate the K_a for the following acids using the given information :

- 0.220M solution of H₃AsO₄ , PH=1.50
- 0.0400M solution of HClO₂ , PH=1.80

12-Calculate the K_a of the following acids using the given information :

- 0.00330M solution of benzoic acid (C₆H₅COOH) , POH =10.70
- 0.100M solution of cyanic acid (HCNO) , POH =11.00
- 0.150M solution of butanoic acid (C₃H₇COOH) , POH =11.18

13-Challenge : calculate the K_a of a 0.0091M solution of an unknown acid (HX) having a POH of 11.32 and identify to acid .

14- the PH of a 0.100 M solution of formic acid (HCOOH) IS 2.38 what is Ka for HCOOH?

Answer (The acid ionization constant for HCOOH is 1.8×10^{-4})

✓ Measuring PH :

- ✓ Indicators are All PH paper that treated with one or more substances
- ✓ Change color in indicators depending on the concentration of H^+ ions in a solution .

Kinds of indicators :

a- Acid – base indicators

- ✓ a litmus paper which used in the lunch lap is an example of a kind of PH paper
- ✓ phenolphthalein
- b- a strip of **PH paper** the color of the paper changes if it is dipped into acidic or a basic solution
 - ✓ to determine the PH , the new color of the paper is compared with standard PH colors on a chart



c- **the PH meter** : provides a more accurate measure of PH

- ✓ when electrodes are placed in a solution
- ✓ the meter gives a direct analog or digital read out of PH

15- complete the following table :

$[H^+]$	$[OH^-]$	PH	POH	Kind of solution
	$1.0 \times 10^{-6} M$			
	$1.0 \times 10^{-4} M$			
$1.0 \times 10^{-9} M$				
$2.5 \times 10^{-2} M$				
	0.000033M			
0.0095M				
		Milk 6.5		
		milk magnesia 10.5		
		lemon 2.37		
		ammonia 11.9		
		tomato 4.5		

16- **challenge** calculate PH and POH for an aqueous solution containing $1.0 \times 10^{-3} \text{ mol}$ of HCl dissolved in 5.0L of solution ?

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Neutralization

MAIN IDEA :

in a neutralization reaction , an acid reacts with a base to produce a salt and water

Item	definition
A neurralization reaction	Is the reaction in which an acid and a base in aqueous solution react to produce a salt and water
A salt	Is an ionic compound made up of a cation from a base and an anion from an acid
titration	Is a method for determining the concentration of a solution by reacting a known volume of that solution with a solution of known concentration
Standard solution (titrant)	Is the titrating solution of known concentration (a buret filled by it)
Equivalence point	Is the point at which moles of H^+ ion from the acid equal moles of OH^- ion from the base
End point	Is the point at which the indicator used in a titration changes color
Acid- base indicators	Are chemical dyes whose colors are affected by acidic and basic solutions
salt hydrolysis	Is the reaction between salt and water which the anions of the dissociated salt accept hydrogen ions from water or the cations of dissociated salt donate hydrogen ions to water
buffers	Are solutions that resist changes in PH when limited amount of acid or base are added
Buffer capacity of solution	The amount of acid or base abuffer solution can absorb without a significant change in PH

✓ reactions between acids and bases :

- ✓ if you were to experience heartburn or digestion to relieve your discomfort you take anti acids like milk of magnesia $Mg(OH)_2$
- ✓ the reaction occurs when magnesium hydroxide $Mg(OH)_2$ the active ingredient in milk of magnesia contacts HCl solution (H^+ , Cl^-) produced by stomach
- ✓ a neutralization reaction occurs
- ✓ neutralization is a double- replacement reaction

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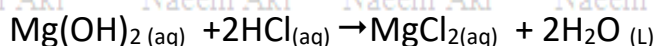
✓ writing neutralization equations:

- ✓ when writing neutralization equations you must know whether all of reactants and products in the solution exist as molecules or as formula units

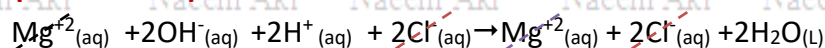
example 1:

write the general ionic equation- net ionic equation for the reaction between $\text{Mg}(\text{OH})_2$ and HCl

solve:

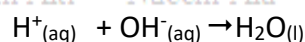
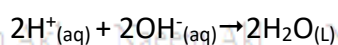


✓ complete ionic equation :

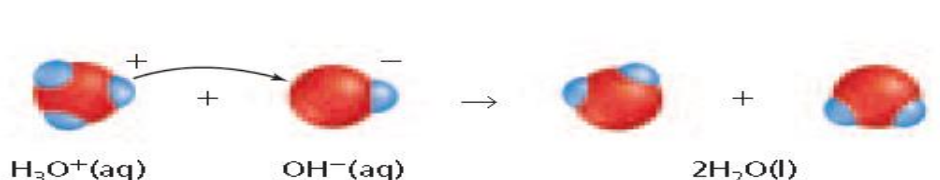


✓ spectator ions Mg^{2+} , 2Cl^{-}

✓ Net ionic equation :



Recall that in aqueous solution a H^{+} ion exists as H_3O^{+} ion so the net ionic equation is:



- ✓ A hydronium ion transfers a hydrogen ion to a hydroxide ion the loss of the hydrogen ion by H_3O^{+} results in water molecule
- ✓ The gain of a hydrogen ion by OH^{-} also results in a water molecule.
- ✓ The net ionic equation of strong acid with strong base :



- ✓ The net ionic equation of weak acid and strong base different from strong acid with strong base

Example : write the net ionic equation between HClO and KOH



Writing the complete ionic equation and the net ionic equation for the neutralization OF:

17- HNO_3 by KOH

18- HCl by NaOH

19- HBr BY CaCO_3

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Acid base titration :

- ✓ The stoichiometry of an acid - base neutralization reaction is the same as that of any other reaction that occurs in solution
- ✓ Stoichiometry provides the basis for a procedure called titration
- ✓ Titration is used to determine the concentration of acidic and basic solution

Titration procedure :

1. Acidic or basic solution of **unknown** concentration is placed in a **beaker**
2. The titrating solution of **known** concentration (standard solution or titrant) placed in a **buret**
3. The electrodes of PH meter are immersed in the solution which is placed in beaker and the initial PH of the solution is recorded
4. Measure volumes of standard solution are added slowly and mixed into a solution in the beaker this process continues until the reaction reaches the **equivalence point**

Example : titration of 50.0ml of 0.100M HCl with 0.100M NaOH

- ✓ The initial **PH** of 0.100M HCl is **1.00**
- ✓ As NaOH is added the solutions **PH increase gradually**
- ✓ When nearly all of the H^+ ions from the acid have been used up the PH increase dramatically with the addition of an exceedingly small volume of NaOH
- ✓ This **abrupt increase** in PH occurs **at the equivalence point** of the titration
- ✓ Beyond the equivalence point the addition of more NaOH again results in a gradual increase in PH
- ✓ Some titrations have equivalence points at PH 7 , others at PH values less than 7 and some have equivalence points at PH values greater than 7
- ✓ A steep rise in the PH of the acid solution indicates that all of the H^+ ions from the acid hve been neutralized by OH^- ions of the base the point at which the curve flexes (at its intersection with the dashed line) is the equivalent point of the titration
- ✓ Bromthymol blue is a good choice for titration of strong acid(HCl) with strong base(NaOH)
- ✓ Phenolphthalein is an indicator that changes color at equivalent point of titration of a weak acid ($HCOOH$) with a strong base(NaOH)



• Acid - base indicators:

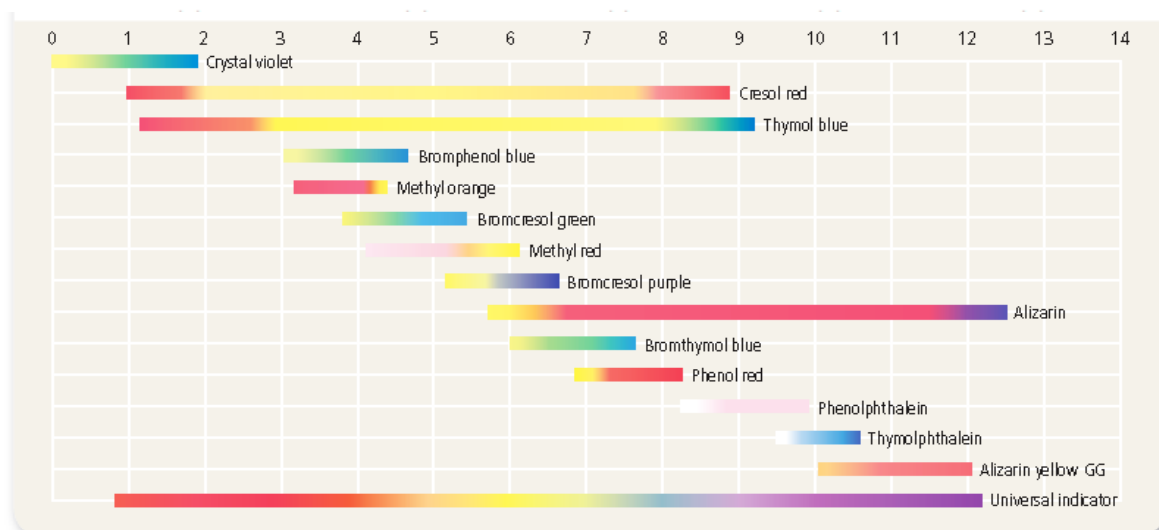
- ✓ Many natural substances act as indicators
- ✓ Chemists often use a chemical dye rather than a PH meter to detect the equivalent point of an acid – base titration

Example:

If you use **lemon juice** in your tea the color of tea gets lighter because tea contains compounds called **polyphenols** that have slightly ionizable hydrogen atoms and therefore are weak acids

Adding acid in the form of lemon juice to a cup of tea depresses the ionization according to le chateliers principle and the color of un- ionized polyphenols becomes more apparent

- ✓ A substance contained in tea is an indicator
- ✓ Most indicators are large molecules that act a weak acids
- ✓ Slight differences in bonding patterns when an indicator molecule is ionized or un ionized account for the color changes



Indicators and titration end point



The buret contains the standard solution (0.1000M NaOH), and the flask contains 25.00 mL HCOOH solution along with a small amount of phenolphthalein indicator.

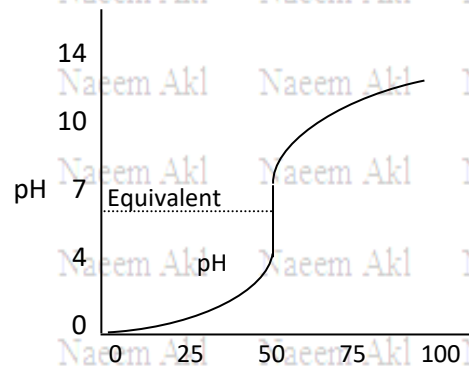
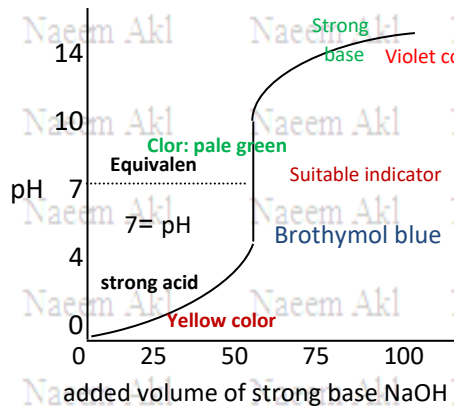


The standard solution is added slowly to the acid solution. The phenolphthalein indicator turns pink, but the color disappears upon mixing, until the end point is reached.

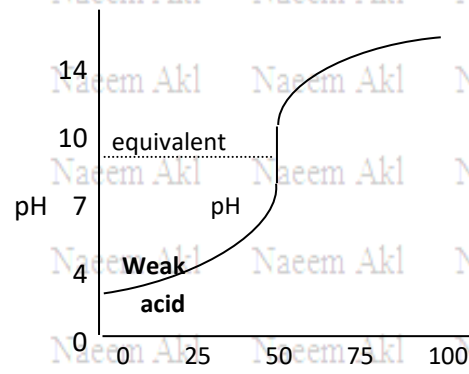
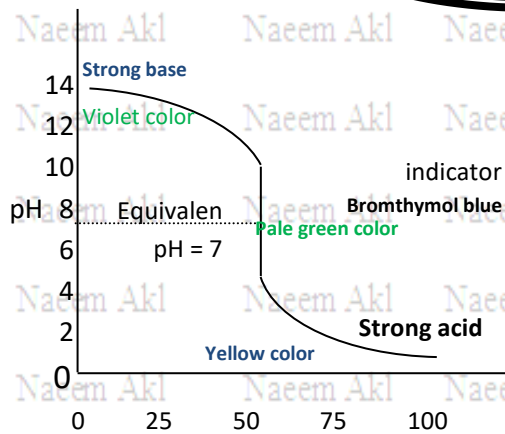


The end point of the titration is marked by a permanent, but very light, pink color. A careful reading of the buret reveals that 18.28 mL 0.1000M NaOH has been added.

• curves of titrations :

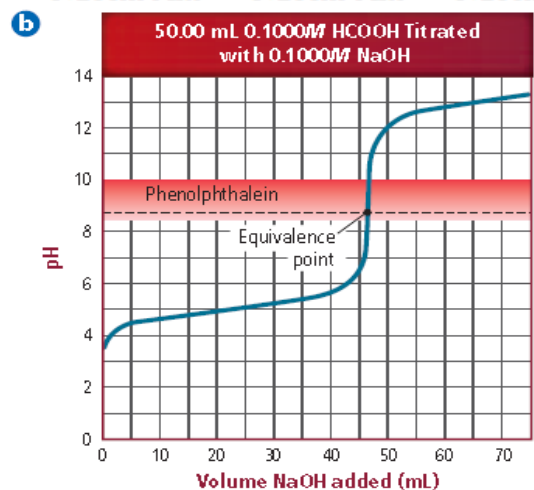
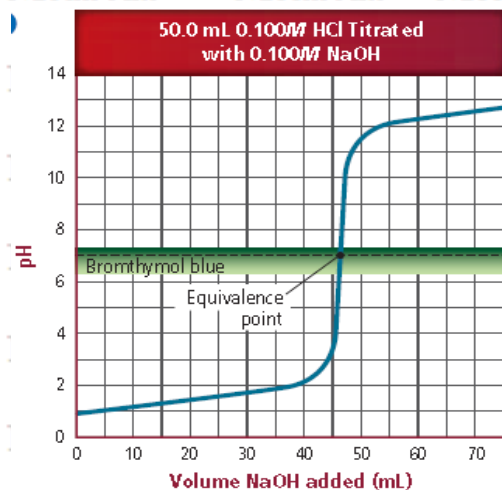


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Added volume of strong acid

Titration of strong acid with strong base

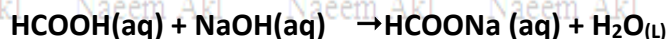


Molarity from titration data:**example 2:**

A volume of 18.28 ml of a standard solution of 0.1000 M NaOH was required to neutralize 25.00ml of a solution of methanoic acid (HCOOH) what is the molarity of the methanoic acid solution ?

Solve :

NaOH القاعدة	HCOOH الحمض
V=18.28	V=25 ml
M=0.1M	M =?
n=1	n=1

Chemical equation

1mol HCOOH يعادل 1mol NaOH

$$V \text{ قاعدة} = 18.28 \text{ ml} \times \frac{1L}{1000 \text{ ml}} = 0.01828L$$

$$\text{mol NaOH} = \frac{0.1 \text{ mol}}{L} \times 0.01828 L = 1.828 \times 10^{-3} \text{ mol NaOH}$$

$$\text{mol HCOOH} = 1.828 \times 10^{-3} \text{ mol NaOH} \times \frac{1 \text{ mol HCOOH}}{1 \text{ mol NaOH}} = 1.828 \times 10^{-3} \text{ mol HCOOH}$$

$$V \text{ حمض} = 25 \text{ ml} \times \frac{1L}{1000 \text{ ml}} = 0.025L$$

$$M \text{ حمض} = \frac{1.828 \times 10^{-3} \text{ mol HCOOH}}{0.025 L \text{ HCOOH}} = 7.312 \times 10^{-2} \text{ mol/L}$$

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20-what is the molarity of a nitric acid solution if 43.33 ml of 0.1000 M KOH solution is needed to neutralize 20.00 ml of acid solution ?

21-what is the concentration of a house hold ammonia cleaning solution if 49.90 ml of 0.5900 M HCl is required to neutralize 25.00 ml of the solution ?

22- what is the molarity of a hydro bromic acid solution if 30.35 ml of 0.1000 M NaOH solution is needed to neutralize 25.00 ml of acid solution ?

23- what is the molarity of a hydro chloric acid solution if 27.4 ml of 0.0154 M Ba(OH)₂ solution is needed to neutralize 20.00 ml of acid solution ?

24- what is the molarity of a sulphoric acid solution if 27.400 ml of 0.0165 M LiOH solution is needed to neutralize 17.60 ml of acid solution ?

25-challenge : how many milliliters of 0.500M NaOH would neutralize 25.00ml of 0.100M H₃PO₄?

تمنيتي لجميع الطلبة والطالبات بالنجاح والتفوق الدائم

تحياتي لـ عبدالله المهيري رئيس مجلس الطلاب

وعلي وراشد المهيري ومعضد

Salt hydrolysis :

- ✓ Prosperities of salts depends on the properties of acid and base which form it
- ✓ Cations (Li^+ , Na^+ , K^+ , Rb^+ , Cs^+ , Ca^{2+} , Sr^{2+} , Ba^{2+}) don't react with water
- ✓ Anions (Cl^- , Br^- , I^- , NO_3^- , ClO_3^- , ClO_4^- , SO_4^{2-}) don't react with water



	Basic salts	Acidic salts	Neutral salts
Strength of acid	Weak (HF)	Strong (HCl)	Strong (HNO ₃)
Strength of base	Strong (KOH)	Weak (NH ₃)	Strong (NaOH)
Example	KF	NH ₄ Cl	NaNO ₃
Salt hydrolysis	$\text{KF (s)} \rightarrow \text{K}^+_{(\text{aq})} + \text{F}^-_{(\text{aq})}$ The K^+ ions do not react with water but the F^- ion is a weak Bronsted- Lowry base Some F^- establish this equilibrium with water $\text{F}^-_{(\text{aq})} + \text{H}_2\text{O (l)} \rightleftharpoons \text{HF}_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$ The production of the OH^- ions make the solution basic	$\text{NH}_4\text{Cl (s)} \rightarrow \text{NH}_4^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})}$ The Cl^- ions do not react with water but the NH_4^+ ion is a weak Bronsted- Lowry acid NH_4^+ ion react with water molecules to establish this equilibrium $\text{NH}_4^+_{(\text{aq})} + \text{H}_2\text{O (l)} \rightleftharpoons \text{NH}_3_{(\text{aq})} + \text{H}_3\text{O}^+_{(\text{aq})}$ The presence of H_3O^+ make the solution acidic	Little or no salt hydrolysis occurs because neither Na^+ nor NO_3^- react with water therefore a solution is neutral
Concentration of H_3O^+ and OH^-	$[\text{H}_3\text{O}^+] < [\text{OH}^-]$	$[\text{H}_3\text{O}^+] > [\text{OH}^-]$	$[\text{H}_3\text{O}^+] = [\text{OH}^-]$
	$K_a < K_b$	$K_a > K_b$	$K_a = K_b$
Bromthymol blue	Turn into blue color	Turn into yellow color	Turn into green color

Practice problems :

26-write equations for the salt hydrolysis reactions occurring when the following salts dissolve in water .classify each acidic , basic , or neutral:

- a- Ammonium nitrate b- potassium sulfate
c-rubidium carbonate d-calcium carbonate

27- challenge : write the equation for the reaction that occurs in a titration of ammonium hydroxide (NH₄OH) with hydrogen bromide (HBr) will the PH at the equivalence point be greater or less than 7 ?

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- **Buffer solutions**

- ✓ **Whats the buffer?**

Are solution that resist changes in PH when limited amount of acid or base are added

- ✓ **Components of buffer solutions**

Is a mixture of :

a Weak acid + its conjugate base (0.1MHF/0.1MNaF)

a weak base + its conjugate acid (0.1MNH₃/0.1M NH₄Cl)

- ✓ **How do buffers work ?**

the mixture of ions and molecules in a buffer solution resists changes in PH by reacting with any H⁺ or OH⁻ added to the buffered solution

Buffer solution (0.1M HF / 0.1M NaF) The NaF Provides a 0.1M concentration of F ⁻ ions $\text{HF}_{(aq)} \rightleftharpoons \text{H}^+_{(aq)} + \text{F}^-_{(aq)}$	
Adding an acid	Adding a base
✓ The acid ionized and increase H⁺ ion in solution ✓ The equilibrium shifts to the left according le chatelier principle ✓ Equilibrium is established again with larger amount of undissociated HF present $\text{H}^+_{(aq)} + \text{F}^-_{(aq)} \longrightarrow \text{HF}_{(aq)}$ ✓ However the PH of solution has changed little because the shift to left consumed most of the added H⁺ ion	✓ the base increase OH⁻ ions in solution ✓ the added OH⁻ react with H⁺ ions to form H₂O ✓ this decrease the concentration of H⁺ ions and equilibrium shifts to the right to replace the H⁺ ions ✓ the PH remains fairly constant because the H⁺ ion concentration has not changed appreciably $\text{HF}_{(aq)} \longrightarrow \text{H}^+_{(aq)} + \text{F}^-_{(aq)}$

- ✓ **Buffer capacity :**

The amount of acid or base abuffer solution can absorb without a significant change in PH

The greater the concentration of the buffering molecules and ions in solution the greater the solutions buffer capacity

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✓ **Choosing buffer :**

A buffer system is most effective when the concentrations of the conjugate acid-base pair are equal or nearly equal

Example

When we mix equal molar amounts of (NaH_2PO_4 و Na_2HPO_4) we obtain a buffer system ($\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$)



What is the PH of such a buffer solution ?

The acid ionization constant expression for the equilibrium can provide the answer

$$K_a = \frac{[\text{HPO}_4^{2-}][\text{H}^+]}{[\text{H}_2\text{PO}_4^-]} = 6.2 \times 10^{-8}$$

- ✓ Because the solution has been made with equal molar amounts of NaH_2PO_4 and Na_2HPO_4

$$[\text{HPO}_4^{2-}] = [\text{H}_2\text{PO}_4^-]$$

$$K_a = \frac{[\text{HPO}_4^{2-}][\text{H}^+]}{[\text{H}_2\text{PO}_4^-]} = 6.2 \times 10^{-8} \quad \checkmark$$

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$$[\text{H}^+] = 6.2 \times 10^{-8}$$

$$\text{pH} = -\log[\text{H}^+] = -\log(6.2 \times 10^{-8}) = 7.21$$

thus when equimolar amounts of each of the components are present

in the ($\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$) buffer system

The system can maintain a PH close to 7.21



Buffer solution $\text{pH} = -\log K_a$

Example

We can make buffer solution when equimolar amount of $\text{HCOOH}/\text{HCOONa}$

Example

We can make buffer solution when equimolar amount of $\text{NH}_3/\text{NH}_4\text{Cl}$ or NH_4NO_3

- ✓ A constant PH is also important in the body (the PH of the blood must be maintained within the range of 7.1 to 7.7)
- ✓ The gastric juices in the stomach must have a PH between 1.6 and 1.8 to promote digestion of certain foods
- ✓ The body maintains PH values within such narrow limits by producing buffers
- ✓ The healthy environment for jellies the PH of the aquarium water at the Monterey bay aquarium (PH of 8.1 to 8.4)

Table 18.7 Buffer Systems with Equimolar Components

Buffer Equilibrium	Conjugate Acid-Base Pair in Buffered Solution	Buffer pH
$\text{HF(aq)} \rightleftharpoons \text{H}^+(\text{aq}) + \text{F}^-(\text{aq})$	HF/F^-	3.20
$\text{CH}_3\text{COOH(aq)} \rightleftharpoons \text{H}^+(\text{aq}) + \text{CH}_3\text{COO}^-(\text{aq})$	$\text{CH}_3\text{COOH}/\text{CH}_3\text{COO}^-$	4.76
$\text{H}_2\text{CO}_3(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{HCO}_3^-(\text{aq})$	$\text{H}_2\text{CO}_3/\text{HCO}_3^-$	6.35
$\text{H}_2\text{PO}_4^-(\text{aq}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{HPO}_4^{2-}(\text{aq})$	$\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$	7.21
$\text{NH}_3(\text{aq}) + \text{H}_2\text{O(l)} \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$	$\text{NH}_4^+/\text{NH}_3$	9.4
$\text{C}_2\text{H}_5\text{NH}_2(\text{aq}) + \text{H}_2\text{O(l)} \rightleftharpoons \text{C}_2\text{H}_5\text{NH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$	$\text{C}_2\text{H}_5\text{NH}_3^+/\text{C}_2\text{H}_5\text{NH}_2$	10.70

How does your blood maintain its PH ?

- ✓ Human blood contains three types of cells
 - ✚ red blood cells : deliver oxygen to every part of the body
 - ✚ white blood cells : fight infections
 - ✚ platelets : aid in clotting when bleeding occurs
- ✓ the critical functions of these cells are impaired if the PH of blood is not maintained within the narrow range of 7.1 to 7.7
- ✓ beyond this range , proteins in the body lose their structures and abilities to function
- ✓ fortunately several buffers maintain the necessary acid/base balance
- ✓ the carbonic acid/ hydrogen carbonate ($\text{H}_2\text{CO}_3 / \text{HCO}_3^-$) buffer is most **important**



- ✓ as acids and bases enter the blood stream as a result of normal activity the bloods buffer systems shift to effectively maintain a healthful PH
- ✓ Depending on the body's metabolic rate and other factors the $\text{H}_2\text{CO}_3/\text{HCO}_3^-$ equilibrium will shift according to le chateliers principle
- ✓ In addition the lungs can alter the rate at which CO_2 is expelled from the body by breathing
- ✓ The kidneys can alter the rate of removal of HCO_3^- ions

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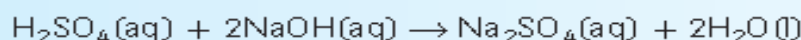
- How many times greater the $[H^+]$ is if the bloods PH changes from PH 7.4 to 7.1 **(more than twice)**
- Suggest a reason why a 20:1 ratio of HCO_3^- to CO_2 in the blood is favorable for maintaining a healthy PH ?
when the body become more active it through the acid in the blood so HCO_3^- make to neutral acid and the reaction produce CO_2
- Predict whether for each situation the PH of the blood will rise or fall and which way the H_2CO_3 / HCO_3^- equilibrium will shift
 - A person will a server stomach virus vomits many times during a 24-h period
Vomiting is acidic it raise the PH , the reaction shifts to right kidneys remove HCO_3^- a person should be calm to keep CO_2 in the body
 - To combat heartburn a person takes too much $NaHCO_3$
PH will increase by increasing the HCO_3^- so the reaction will shift to left producing CO_2 kidneys will respond by removing HCO_3^- and a person can breath quickly to release CO_2

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Problem-Solving Strategy

Calculating Molarity

The balanced equation for a titration reaction is the key to calculating the unknown molarity. For example, sulfuric acid is titrated with sodium hydroxide according to this equation.



- Calculate the moles of NaOH in the standard from the titration data: molarity of the base (M_B) and the volume of the base (V_B).
 $(M_B)(V_B) = (mol/L)(L) = mol \text{ NaOH in standard}$
- From the equation, you know that the mole ratio of NaOH to H_2SO_4 is 2:1. Two moles of NaOH are required to neutralize 1 mol of H_2SO_4 .
 $mol \text{ } H_2SO_4 \text{ titrated} = mol \text{ NaOH in standard} \times \frac{1 \text{ mol } H_2SO_4}{2 \text{ mol NaOH}}$
- M_A represents the molarity of the acid and V_A represents the volume of the acid in liters. $M_A = \frac{mol \text{ } H_2SO_4 \text{ titrated}}{V_A}$

Apply this strategy as you study Example Problem 18.6.

وفي الختام

أهدي هذا العمل المتواضع الي موجهي الكيمياء

أ. عادل الجيار - أ. محمد ابو عبيد

داعيا المولي عز وجل لهما بالتوفيق والسداد لما قدموه لنا من علم وخلق

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