أوراق عمل الوحدة الرابعة Bases and Acids الأحماض والأسس منهج انسباير

المزيد من الملفات بحسب الصف الثاني عشر المتقدم والمادة كيمياء في الفصل الثاني	
5	حل مراجعة نهائية هامة وفق الهيكل الوزاري

Chemistry Worksheets

Unit 4

Second Term 2024-25

12ADV

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الحل على موقع الأكاديمية

+971 50 620 6080
https://www.manasra.academy
تقديم واعداد الأستاذ: عبدالرحيم قدومي

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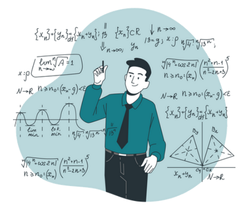
حصص مسجلة لشرح كامل مادة الفصل الثاني 	امتحانات الكترونية محاكية لامتحانات الوزارة 	حصص مسجلة لمراجعة الهيكل وأسئلة السنوات 
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CHAPTER 4 / Acids and Bases

Section 1: Introduction to Acids and Bases

Properties of Acids and Bases:

The Acid	Methanoic Acid (Formic Acid)	Carbonic and Phosphoric acids	Hydrochloric acid	Cetric and Ascorbic acids	Acetic acid
Its presence	In Ants	In carbonated beverages	In stomach	In Lemon and Grapefruit	In vinegar

- Uses of bases:
  - Sodium hydroxide uses in making of soap.
  - Magnesium hydroxide uses in making of antiacid tablets.

Reactivity series
.
.
.
H
Cu
Ag
Hg
Pt
Au

Physical properties:

Acids	Bases
1- Acidic solutions taste sour	1- Basic solutions taste bitter and feel slippery
2- Acidic solutions are electricity conductors (produce ions)	2- Basic solutions are electricity conductors (produce ions)

Chemical Properties:

Acids	Bases
1- Aqueous solutions of acids cause blue litmus paper to turn red	1- Aqueous solutions of bases cause red litmus paper to turn blue
2- React with active metals to produce hydrogen gas: $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$	2- Not react with metals.
3- React with metal carbonates and hydrogen carbonates to produce CO_2 $NaHCO_3(s) + HC_2H_3O_2(aq) \rightarrow NaC_2H_3O_2(aq) + H_2O(l) + CO_2(g)$	3- Not react with metal carponates

Question: How did the geologists identify rocks as limestone ($CaCO_3$)? Write equation.

PRACTICE Problems

1. Write balanced equations for the reactions between the following.
   a. aluminum and sulfuric acid
   b. calcium carbonate and hydrobromic acid
2. **Challenge** Write the net ionic equation for the reaction in Question 1b.

Hydronium and Hydroxide Ions:

- **All water solutions** contain hydrogen ions (H^+) and hydroxide ions (OH^-).
  - **An acidic solution** contains more hydrogen ions than hydroxide ions.
  - **A basic solution** contains more hydroxide ions than hydrogen.
  - **A neutral solution** contains equal concentrations of H^+ and OH^-.
- **Self – ionization of water:** water molecules react to form a hydronium ion (H_3O^+) and a hydroxide ion.

$$H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)} \quad \text{or} \quad H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

Water molecules — Hydronium ion — Hydroxide ion

- **The hydronium ion:** is a hydrogen ion which has a water molecule attached to it by a covalent bond.

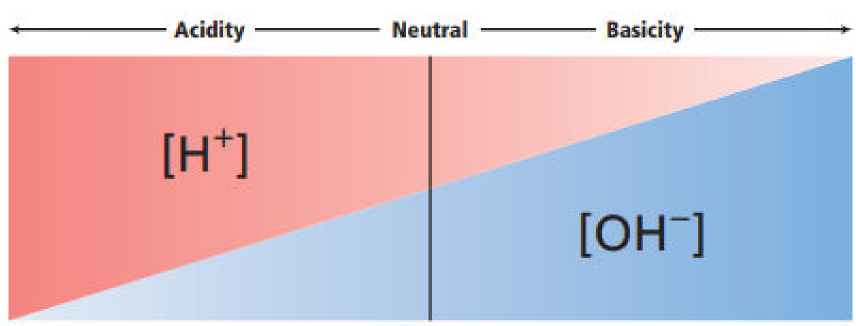


Figure 18.3 Note how [H+] and [OH−] change simultaneously. As [H+] decreases to the right, [OH−] increases to the right.
Identify *the point in the diagram at which the two ion concentrations are equal.*

The Arrhenius Model:

- **An acid**: is a substance that contains hydrogen and ionizes to produce hydrogen ions in aqueous solution.

$$HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$$

- **A base**:is a substance that contains a hydroxide group and dissociates to produce a hydroxide ion in aqueous solution.

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

- **Arrhenius model shortcoming**:
  - ammonia (NH_3) and sodium carbonate (Na_2CO_3) do not contain a hydroxide group, yet both substances produce hydroxide ions insolution.
  - Sodium carbonate is the compound that causes the alkalinity of Lake Natron, Tanzania.

The Brønsted-Lowry Model:

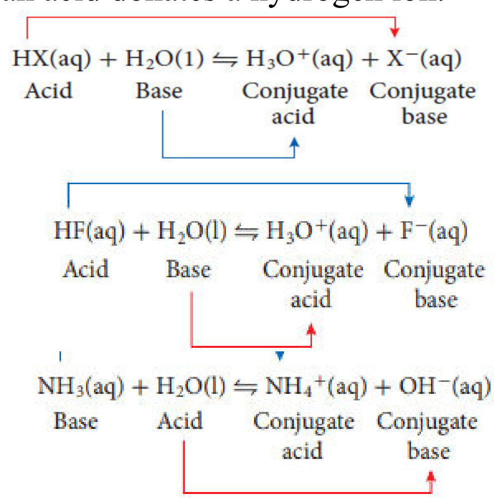
- An acid is a hydrogen-ion donor.
- A base is a hydrogenion acceptor.

QUESTION: determine the Bronsted-Lowry acid and base in the following reactions:

The reaction	The acid	The base
$HCl + NH_3 \rightarrow NH_4^+ + Cl^-$		
$HCl + H_2O \rightarrow H_3O^+ + Cl^-$		
$H_2O + NH_3 \rightarrow NH_4^+ + OH^-$		
$HCO_3^- + H_2O \rightarrow H_2CO_3 + OH^-$		

Hydrogen ion donors and acceptors:

- **A conjugate acid**: is the species produced when a base accepts a hydrogen ion.
- **A conjugate base**: is the species that results when an acid donates a hydrogen ion.
- **A conjugate acid base pair**: consists of two substances related to each other by thedonating and accepting of a single hydrogen ion.
  - Hydrogen fluoride is used to manufacture a variety of fluorine containing compounds, such as the nonstick coating on the kitchenware.



- **Complete the following table:**

The compound or ion	The conjugate acid	The compound or ion	The conjugate base
H_2O		HF	
CO_3^{-2}		H_2	
NH_3		H_2O	
H^-		OH^-	

❖ **Amphoteric substance:** A **substance that can act as both acids and bases , like water and any anion that contain hydrogen like HCO_3^- .**

Question: Explain how the ion HCO_3^- can be both an acid and a base (**amphoteric substance**).

PRACTICE Problems **Extra Practice** Page 989 and glencoe.com

3. Identify the conjugate acid-base pairs in each reaction.
   a. $NH_4^+(aq) + OH^-(aq) \rightleftharpoons NH_3(aq) + H_2O(l)$
   b. $HBr(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Br^-(aq)$
   c. $CO_3^{2-}(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$
4. **Challenge** The products of an acid-base reaction are H_3O^+ and SO_4^{2-}. Write a balanced equation for the reaction and identify the conjugate acid-base pairs.

Monoprotic and Polyprotic Acids:

❖ Monoprotic acid : An acid that can donate only onehydrogen ion like HCl,CH_3COOH

❖ Polyprotic acid: any acid that has more than one ionizable hydrogen atom.

- Diprotic acid: An acid that contain two ionizable hydrogen atoms per molecule like H_2SO_4 :

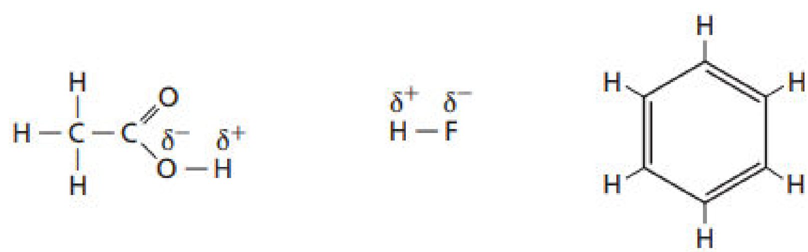
- A triprotic acid: An acid with three hydrogen ions to donate like H_3PO_4 , H_3BO_3.

Question: write the ionizations steps of phosphoric acid:

Ionizable hydrogen atoms:

- The difference in electronegativity makes the bond between oxygen and hydrogen polar, and weak , then it will be ionize easily.
- the hydrogen atoms in benzene are each bonded to acarbon atom. Carbon atoms have about the same electronegativity as hydrogen. These bonds are nonpolar, so benzene is not an acid.

Figure 18.9 Whether a hydrogen is ionizable depends on the polarity of its bond. In acetic acid, oxygen is more electronegative than hydrogen. The bond between oxygen and hydrogen is polar, so the hydrogen atom can ionize in solution. In hydrogen fluoride, fluorine is highly electronegative, so HF is an acid in solution. In benzene, there is little electronegativity difference between the carbon and hydrogen atoms, so benzene is not an acid.

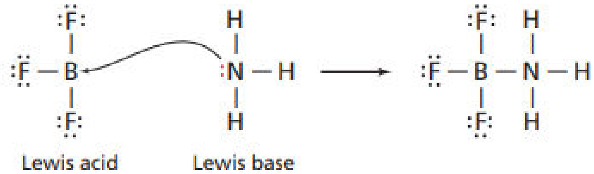


Acetic acid   Hydrogen fluoride   Benzene

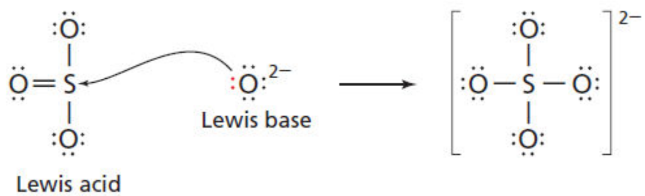
Table 18.1	**Some Common Acids and Their Conjugate Bases**		
Acid		Congugate Base	
Name	Formula	Name	Formula
Hydrochloric acid	HCl	Chloride ion	Cl^-
Nitric acid	HNO_3	Nitrate ion	NO_3^-
Sulfuric acid	H_2SO_4	Hydrogen sulfate ion	HSO_4^-
Hydrogen sulfate ion	HSO_4^-	Sulfate ion	SO_4^{2-}
Hydrofluoric acid	HF	Fluoride ion	F^-
Hydrocyanic acid	HCN	Cyanide	CN^-
Acetic acid	$HC_2H_3O_2$	Acetate ion	$C_2H_3O_2^-$
Phosphoric acid	H_3PO_4	Dihydrogen phosphate ion	$H_2PO_4^-$
Dihydrogen phosphate ion	$H_2PO_4^-$	Hydrogen phosphate ion	HPO_4^{2-}
Hydrogen phosphate ion	HPO_4^{2-}	Phosphate ion	PO_4^{3-}
Carbonic acid	H_2CO_3	Hydrogen carbonate ion	HCO_3^-
Hydrogen carbonate ion	HCO_3^-	Carbonate ion	CO_3^{2-}

The Lewis Model:(more general model)

- **Lewis acid** :is an ion or molecule with a vacant atomic orbital that can accept (share) an electron pair , like : BF_3 , BCl_3 , AlF_3 , $AlCl_3$ or any positive ion H^+ , Ag^+ .
- **Lewis base** :an ion or molecule with a lone electron pair that it can donate (share), like : NH_3, H_2O or any negative ion Cl^- , F^- .



$$SO_3(g) + MgO(s) \rightarrow MgSO_4(s)$$



- The reaction of SO_3 and MgO is important because it produces magnesium sulfate, a salt that forms the heptahydrate known as ……………………………………...
- Epsom salt has many uses, including …………………………………… and acting as a plant…………………. .
- The reaction to form magnesium sulfate also has environmental applications:
  1- MgO is injected into the flue gases of coal-fired power plants.
  2- It reacts with and removes SO_3 .
  (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) .

Model	Acid Definition	Base Definition
Arrhenius	H^+ producer	OH^- producer
Brønsted-Lowry	H^+ donor	H^+ acceptor
Lewis	electron-pair acceptor	electron-pair donor

H.W : SOLVE questions 55 - 64 .

SECTION 1 REVIEW

5. A Lewis acid is an electron pair acceptor. A Lewis base is an electron pair donor. A Lewis acid cannot have an ionizable hydrogen ion or hydroxide ion to qualify as an Arrhenius acid or base. A Lewis acid cannot have a hydrogen ion to donate, therefore it could not qualify as a Brønsted-Lowry acid. However, all Lewis bases are Brønsted-Lowry bases because they can accept a hydrogen ion.
6. Physical properties: Acids taste sour and conduct electricity. Bases taste bitter, feel slippery, and conduct electricity. Chemical properties: Acids react with some metals to produce hydrogen gas. Acids turn blue litmus red. Bases react with acids and turn red litmus blue.
7. In an acidic solution, $[H^+] > [OH^-]$; in a neutral solution, $[H^+] = [OH^-]$; in a basic solution, $[H^+] < [OH^-]$.
8. Only compounds that have one or more ionizable hydrogen atom are Arrhenius acids.
9. HNO_2 (acid) and NO_2^- (conjugate base), H_2O (base) and H_3O^+ (conjugate acid)
10. Phosphorus in PCl_3 has three electrons, which it shares with three chlorines, and an unshared pair of electrons. The unshared pair of electrons can act as a Lewis base.
11. the two hydrogen atoms attached to oxygen atoms

For each description below, write *acid* if it tells about a property of an acid or *base* if it tells about a property of a base. If the property does not apply to either an acid or a base, write *neither*. If it applies to both an acid and a base, write *both*.

______________ **1.** Can turn litmus paper a different color

______________ **2.** Reacts with certain metals

______________ **3.** Contains more hydrogen ions than hydroxide ions

______________ **4.** Feels slippery

______________ **5.** Reacts with carbonates

______________ **6.** Feels rough

______________ **7.** Contains equal numbers of hydrogen and hydroxide ions

______________ **8.** Tastes bitter

______________ **9.** Tastes sour

Arrhenius	Brønsted-Lowry	conjugate acid
conjugate base	hydrogen	hydroxide

The **(10)**____________________ model of acids and bases states that an acid contains the element **(11)**____________________ and forms ions of this element when it is dissolved in water. A base contains the **(12)**____________________ group and dissociates to produce **(13)**____________________ ions in aqueous solution.

According to the **(14)**____________________ model, an acid donates **(15)**____________________ ions, and a base accepts **(16)**____________________ ions. According to this model, in an acid-base reaction, each acid has a **(17)**____________________, and each base has a **(18)**____________________.

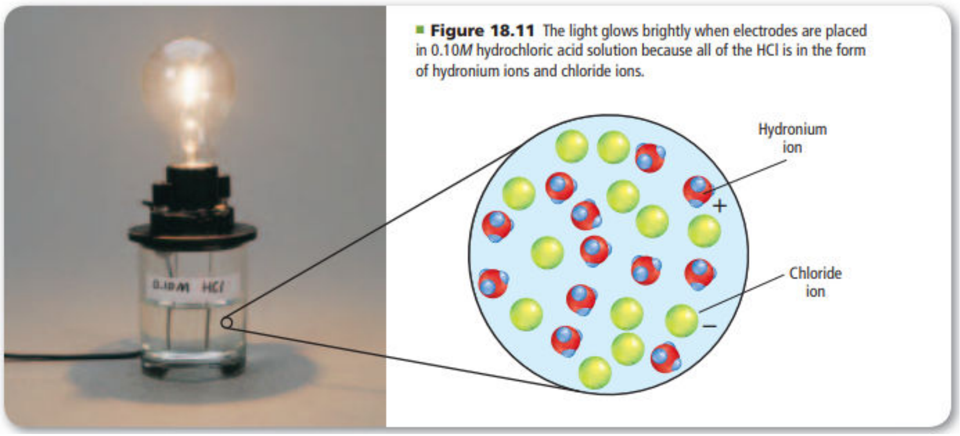
Section 2 : Strengths of of Acids and Bases:

<u>Strength of Acids:</u>

❖ **Degree of ionization:**

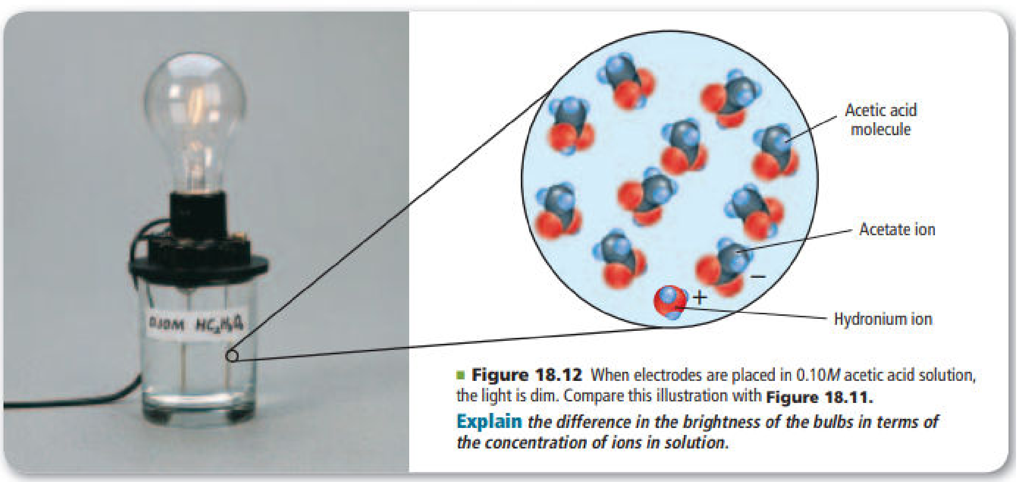
- Strong Acids: completely ionized (strong electrolytes) and good conductors , because strong acids produce the maximum number of ion.

$$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$



■ **Figure 18.11** The light glows brightly when electrodes are placed in 0.10*M* hydrochloric acid solution because all of the HCl is in the form of hydronium ions and chloride ions.

- Weak Acids: partially ionized (weak electrolytes) and weak conductors ,because weak acids produce fewer ions.



■ **Figure 18.12** When electrodes are placed in 0.10*M* acetic acid solution, the light is dim. Compare this illustration with **Figure 18.11**. **Explain** *the difference in the brightness of the bulbs in terms of the concentration of ions in solution.*

Concepts In Motion

Table 18.3 Ionization Equations

Interactive Table Explore ionization equations at glencoe.com.

Strong Acids		Weak Acids	
Name	**Ionization Equation**	**Name**	**Ionization Equations**
Hydrochloric	$HCl \rightarrow H^+ + Cl^-$	Hydrofluoric	$HF \rightleftharpoons H^+ + F^-$
Hydroiodic	$HI \rightarrow H^+ + I^-$	Acetic	$HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-$
Perchloric	$HClO_4 \rightarrow H^+ + ClO_4^-$	Hydrosulfuric	$H_2S \rightleftharpoons H^+ + HS^-$
Nitric	$HNO_3 \rightarrow H^+ + NO_3^-$	Carbonic	$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$
Sulfuric	$H_2SO_4 \rightarrow H^+ + HSO_4^-$	Hypochlorous	$HClO \rightleftharpoons H^+ + ClO^-$

❖ Acid strength and the Brønsted-Lowry model:

- HX represents a strong acid and its conjugate base is weak.
- H_2O is a stronger base (in the forward reaction) than is the conjugate base X^- (in the reverse reaction).
- the ionization equilibrium lies almost completely to **the right** because the base H_2O has a much greater attraction for the H^+ ion than does the base X^-.

$$HX(aq) + H_2O(l) \rightarrow H_3O^+(aq) + X^-(aq)$$

Acid | Base | Conjugate acid | Conjugate base

- The ionization equilibrium for a weak acid lies far to the **<u>left</u>** because the conjugate base Y - has a greater attraction for the H^+ ion than does the base H_2O.
- The conjugate base Y^- (in the reverse reaction) is stronger than the base H_2O (in the forward reaction)

$$HY(aq) + H_2O(l) \leftrightharpoons H_3O^+(aq) + Y^-(aq)$$

Acid | Base | Conjugate acid | Conjugate base

Acid ionization constants:

- The equilibrium constant, K_{eq} , provides a quantitative measure of the degree of ionization of the acid acid strength**.**
  - Consider hydrocyanic acid (HCN), also known as **<u>prussic acid</u>** which is used in : **dying**, **engraving**, and **tempering** steel.

$$HCN(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CN^-(aq)$$
$$K_{eq} = \frac{[H_3O^+][CN^-]}{[HCN][H_2O]}$$

$$K_{eq}[H_2O] = K_a = \frac{[H_3O^+][CN^-]}{[HCN]} = 6.2 \times 10^{-10}$$

❖ The acid ionization constant: is the value of the equilibrium constant expression for the ionization of a weak acid.
❖ The value of K_a indicates whether reactants or products are favored at equilibrium

PRACTICE Problems — **Extra Practice** Page 989 and **glencoe.com**

12. Write ionization equations and acid ionization constant expressions for each acid.
a. $HClO_2$ **b.** HNO_2 **c.** HIO

13. Write the first and second ionization equations for H_2SeO_3.

14. Challenge Given the expression $K_a = \frac{[AsO_4^{3-}][H_3O^+]}{[HCN]}$, write the balanced equation for the corresponding reaction.

Table 18.4 Ionization Constants for Weak Acids

Acid	Ionization Equation	K_a (298 K)
Hydrosulfuric, first ionization	$H_2S \rightleftharpoons H^+ + HS^-$	8.9×10^{-8}
Hydrosulfuric, second ionization	$HS^- \rightleftharpoons H^+ + S^{2-}$	1×10^{-19}
Hydrofluoric	$HF \rightleftharpoons H^+ + F^-$	6.3×10^{-4}
Hydrocyanic	$HCN \rightleftharpoons H^+ + CN^-$	6.2×10^{-10}
Acetic	$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$	1.8×10^{-5}
Carbonic, first ionization	$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$	4.5×10^{-7}
Carbonic, second ionization	$HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$	4.7×10^{-11}

Strengths of Bases:

- **Strong base** : a base that dissociates entirely into metal ions and hydroxide ions

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

- calcium hydroxide and other slightly soluble metallic hydroxides are considered strong bases because all of the compound that dissolves is completely dissociated.

$$Ca(OH)_2(s) \rightleftharpoons Ca^{2+}(aq) + 2OH^-(aq) \quad K_{sp} = 6.5 \times 10^{-6}$$

- **A weak base** ionizes only partially in dilute aqueous solution.

$$CH_3NH_2(aq) + H_2O(l) \rightarrow CH_3NH_3^+(aq) + OH^-(aq)$$

Acid Base Conjugate acid Conjugate base

Table 18.5 Dissociation Equations for Strong Bases

$NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$
$KOH(s) \rightarrow K^+(aq) + OH^-(aq)$
$RbOH(s) \rightarrow Rb^+(aq) + OH^-(aq)$
$CsOH(s) \rightarrow Cs^+(aq) + OH^-(aq)$
$Ca(OH)_2(s) \rightarrow Ca^{2+}(aq) + 2OH^-(aq)$
$Ba(OH)_2(s) \rightarrow Ba^{2+}(aq) + 2OH^-(aq)$

Table 18.6 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}

- **Base ionization constants K_b:**
  Is the value of the equilibrium constant expression for the ionization of a base.

PRACTICE Problems — **Extra Practice** Page 989 and **glencoe.com**

15. Write ionization equations and base ionization constant expressions for the following bases.

a. hexylamine ($C_6H_{13}NH_2$)
b. propylamine ($C_3H_7NH_2$)
c. carbonate ion (CO_3^{2-})
d. hydrogen sulfite ion (HSO_3^-)

16. 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^-.

SECTION 2 REVIEW

17. The solution of HI contains only H_3O^+ and I^- ions and water molecules. The solution of HCOOH contains H_3O^+ and $HCOO^-$ ions, and HCOOH and H_2O molecules.

18. The stronger the acid is, the weaker its conjugate base. The weaker the acid is, the stronger its conjugate base.

19. a. acid: HCOOH; conjugate base: $HCOO^-$; base: H_2O; conjugate acid: H_3O^+

b. acid: H_2O; conjugate base: OH^-; base: NH_3; conjugate acid: NH_4^+

20. The size of the K_b indicates that aniline is a weak base.

21. HS^-, HCO_3^-, HCN, H_2S, H_2CO_3, CH_3COOH, HF

H.W: Solve the question 65 – 74 .

Circle the letter of the choice that best completes the statement or answers the question.

1. Acid A and acid B are of equal concentration and are tested with a conductivity apparatus. When the electrodes are placed in acid A, the bulb glows dimly. When they are placed in acid B, the bulb glows more brightly. Which of the following is true?

a. Acid A is stronger than acid B.

b. Acid B is stronger than acid A.

c. Acid A and acid B are of equal strength.

d. No comparison of strength can be made from the results.

2. A chemical equation for the ionization of an acid uses a single arrow to the right ($\rightarrow$) to separate the reactant and product sides of the equation. Which of the following is true?

a. The arrow does not indicate relative strength.

b. The ionizing acid is half ionized.

c. The ionizing acid is strong.

d. The ionizing acid is weak.

3. Sulfuric acid is a strong acid. What is true about its conjugate base?

a. Its conjugate base is amphoteric.

b. Its conjugate base is strong.

c. Its conjugate base is weak.

d. No conclusion can be made regarding the strength of the conjugate base.

4. In solution, a weak acid produces

a. a mixture of molecules and ions.

b. all ions.

c. all molecules.

d. anions, but no hydronium ions.

5. Why are K_a values all small numbers?

a. The concentration of water does not affect the ionization.

b. The equilibrium is not stable.

c. The solutions contain a high concentration of ions.

d. The solutions contain a high concentration of un-ionized acid molecules.

6. Which of the following dissociates entirely into metal ions and hydroxide ions in solution?

a. a strong acid **b.** a strong base **c.** a weak acid **d.** a weak base

7. In general, compounds formed from active metals, and hydroxide ions are

a. strong acids. **b.** strong bases. **c.** weak acids. **d.** weak bases.

Section 3 : Hydrogen Ions and PH:

Ion Product Constant for Water:

- The self ionization equation of water is : $H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$

The Ion Product of Water

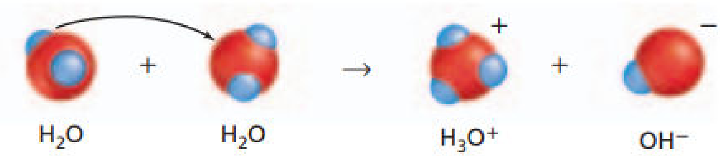
$$K_w = [H^+][OH^-]$$

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

- The ion product constant for water K_w : is the value of the equilibrium constant expression for the self-ionization of water.

Figure 18.13 In the self-ionization of water, one water molecule acts as an acid, and the other acts as a base.



- Experiments show that in pure water at 298 K, [H +] and [O H -] are both equal to 1.0×10^{-7}M.
- Therefore, at 298 K, the value of Kw is 1.0×10^{-14} .

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

- **Explain** why K_w does not change when the concentration of hydrogen ions increases. (in terms of Le Châtelier's principle**)**

……

……

Question: Complete the following tables:

	$[H^+]$	Type of the solution (Acidic , Basic or neutral)	$[OH^-]$	Type of the solution (Acidic , Basic or neutral)
	1.0×10^{-4}M		1.0×10^{-4}M	
	1.0×10^{-11} M		1.0×10^{-11} M	
	1.0×10^{-7} M		**1.0×10^{-7} M**	

- Calculate the **$[H^+]$ or $[OH^-]$**

$[OH^-]$	$[H^+]$
	1×10^{-5} M
	1×10^{-13} M
1×10^{-3} M	
	4×10^{-5} M

- **Challenge** Calculate the number of H^+ ions and the number of OH^- ions in 300 mL of pure water at 298 K.

----

pH and pOH:

- **pH of a solution: is the negative logarithm of the hydrogen 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.

  - The logarithmic nature of the pH scale means that a change of one pH unit represents a tenfold change in ion concentration.
  - A solution having a pH of 3.0 has ten times the hydrogen ion concentration of a solution with a pH of 4.0.

- **pOH of a solution: is the negative logarithm of the hydroxide 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.

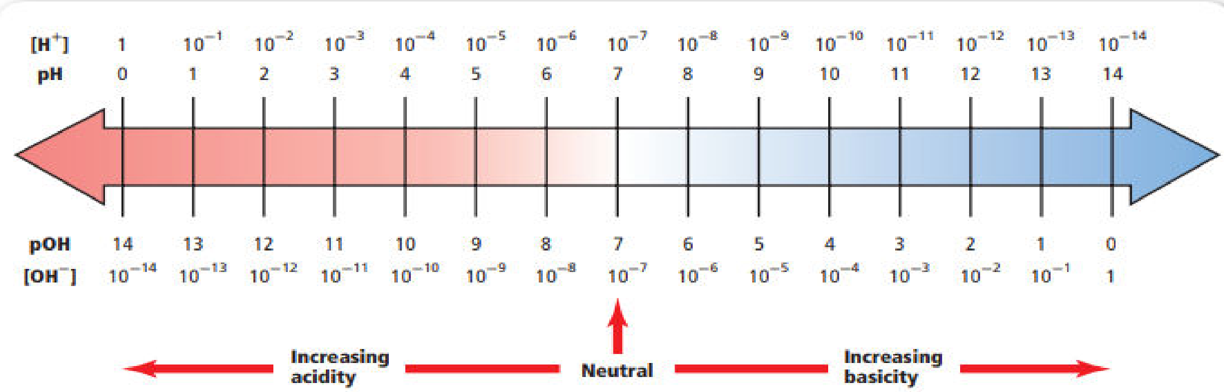


Figure 18.15 Study this diagram to sharpen your understanding of pH and pOH. Note that at each vertical position, the sum of pH (above the arrow) and pOH (below the arrow) equals 14. Also note that at every position, the product of $[H^+]$ and $[OH^-]$ equals 10^{-14}.

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24. Calculate the pH of solutions having the following ion concentrations at 298 K.

a. $[H^+] = 1.0 \times 10^{-2} M$ **b.** $[H^+] = 3.0 \times 10^{-6} M$

25. Calculate the pH of aqueous solutions with the following $[H^+]$ at 298 K.

a. $[H^+] = 0.0055 M$ **b.** $[H^+] = 0.000084 M$

26. Challenge Calculate the pH of a solution having $[OH^-] = 8.2 \times 10^{-6} M$.

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27. Calculate the pH and pOH of aqueous solutions with the following concentrations at 298 K.

a. $[OH^-] = 1.0 \times 10^{-6} M$

b. $[OH^-] = 6.5 \times 10^{-4} M$

c. $[H^+] = 3.6 \times 10^{-9} M$

d. $[H^+] = 2.5 \times 10^{-2} M$

28. Calculate the pH and pOH of aqueous solutions with the following concentration at 298 K.

a. $[OH^-] = 0.000033 M$

b. $[H^+] = 0.0095 M$

29. Challenge Calculate pH and pOH for an aqueous solution containing 1.0×10^{-3} mol of HCl dissolved in 5.0 L of solution.

Calculating ion concentrations from pH:

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30. Calculate $[H^+]$ and $[OH^-]$ in each of the following solutions.

a. Milk, pH = 6.50.
b. Lemon juice, pH = 2.37
c. Milk of magnesia, pH = 10.50
d. Household ammonia, pH = 11.90

31. Challenge Calculate the $[H^+]$ and $[OH^-]$ in a sample of seawater with a pOH = 5.60.

Molarity and the pH of strong acids:

- For all strong monoprotic acids, the concentration of the acid is the concentration of H^+ ions. Thus, you can use the molarity of the acid to calculate pH .

$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

Molarity and the pH of strong bases:

- One formula unit of NaOH produces one OH^- ion. Thus, the concentration of the OH^- ions is the same as the molarity of the solution, 0.1M .
- Some strong bases, such as calcium hydroxide $Ca(OH)_2$, contain two or more OH^- ions in each formula unit. The concentration of OH^- ion in a solution of $Ca(OH)_2$ is twice the molarity of the ionic compound
- Remember that weak acids and weak bases are only partially ionized. Therefore, you must use K_a and K_b values to determine the concentrations of H^+ and OH^- ions in solutions of weak acids and bases.

Calculating Ka from pH:

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32. Calculate the K_a for the following acids using the given information.
a. 0.220*M* solution of H_3AsO_4, pH = 1.50 **b.** 0.0400*M* solution of $HClO_2$, pH = 1.80

33. Calculate the K_a of the following acids using the given information.
a. 0.00330*M* solution of benzoic acid (C_6H_5COOH), pOH = 10.70
b. 0.100*M* solution of cyanic acid (HCNO), pOH = 11.00
c. 0.150*M* solution of butanoic acid (C_3H_7COOH), pOH = 11.18

34. Challenge Calculate the K_a of a 0.0091*M* solution of an unknown acid (HX) having a pOH of 11.32. Use **Table 18.4** to identify to acide.

Measuring pH:

- indicator papers to measure the pH of a solution.
- All pH paper is treated with one or more substances called indicators that change color depending on the concentration of hydrogen ions in a solution.
- pH paper:To determine the pH, the new color of the paper is compared with standard pH colors on a chart.
- The pH meter: provides a more accurate measure of pH.

SECTION 3 REVIEW

35. The sum of pH and pOH is 14.00. If a solution is acidic, its pH is less than 7.00. Therefore, pOH must be greater than 7.00.

36. Subtract the pOH from 14.00.

37. If one ion concentration is known, the other can be calculated using the K_w expression.

38. The increase in OH^- ion from the drop of NaOH shifts the self-ionization of water toward the left and increases the amount of undissociated water molecules. $[OH^-]$ increases and $[H^+]$ decreases.

39. the pH, pOH, or $[H^+]$ and the initial concentration of the acid; or K_b

40. $[H^+] = 3.2 \times 10^{-5} M$, $[OH^-] = 3.2 \times 10^{-10} M$

41. pH = 5.00

42. **a.** pH = 0.00 **c.** pH = 14.00
b. pH = 1.30 **d.** pH = 9.68

43. As the solution becomes more acidic, $[H^+]$ increases from 10^{-7} to 1, $[OH^-]$ decreases from 10^{-7} to 10^{-14}, pH changes from 7 to 0 and pOH changes from 7 to 14. As a neutral solution becomes more basic, $[H^+]$ decreases from 10^{-7} to 10^{-14}, $[OH^-]$ increases from 10^{-7} to 1, pH changes from 7 to 14 and pOH changes from 7 to 0.

H.W: Solve the questions 75 – 84 .

Answer the following questions.

1. Write the simplest form of the chemical equation for the self-ionization of water.

____

2. Write the equilibrium constant expression, K_{eq}, for this equation.

____

3. Write the expression for the equilibrium constant for water, K_w.

____

4. Why can the concentration of water be ignored in the equilibrium expression for water?

____

5. What is the numerical value of K_w at 298 K?

____

6. In solution, if the hydroxide ion concentration increases, what happens to the hydrogen ion concentration?

____

7. If the concentration of hydroxide ions in solution is 1.0×10^{-6}, what is the hydrogen ion concentration?

____

8. Is the solution in question 7 acidic, basic, or neutral? Explain.

____

In the space at the left, write *true* if the statement is true; if the statement is false, change the italicized word or number to make it true.

______________ **9.** The pH of a solution is the negative logarithm of its *hydroxide* ion concentration.

______________ **10.** Values for pH range from *0 to 14*.

______________ **11.** Stomach contents can have a pH of 2, which means that they are *basic*.

______________ **12.** The hydrogen ion concentration in a solution with a pH of 3 is *two* times greater than the hydrogen ion concentration in a solution with a pH of 5.

______________ **13.** The pH of a neutral solution at room temperature *equals* the pOH of the solution.

______________ **14.** If the pH of a solution is 3, its pOH is *10*.

______________ **15.** The pH of a solution with a $[H^+]$ of 1×10^{-8} is *8*.

______________ **16.** The pH of a solution with a $[OH^-]$ of 1×10^{-6} is *6*.

17. What is the pH of a $4.3 \times 10^{-2}M$ HCl solution? HCl is a strong acid.

18. Calculate the pH of a $5.2 \times 10^{-3}M$ H_2SO_4 solution? H_2SO_4 is a strong acid.

19. What is the pH of a $2.5 \times 10^{-5}M$ NaOH solution? NaOH is a strong base.

20. Calculate the pH of a $3.6 \times 10^{-6}M$ $Ca(OH)_2$ solution. $Ca(OH)_2$ is a strong base.

Section4 :Neutralization

- **A neutralization reaction:** is a reaction in which an acid and a base in an aqueous solution react to produce a salt and water.
- A salt: is an ioniccompound made up of a cation from a base and an anion from an acid.
- **Write** the net ionic equation for the following neutralization reaction:

$$Mg(OH)_2(aq) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$$
$$\text{Base} + \text{Acid} \rightarrow \text{Salt} + \text{Water}$$

Acid-base titration:

- **Titration** :is a method for determining the concentration of a solutionby reacting a known volume of that solution with a solution of known concentration.(if one of them is acid the other is base)

Titration procedure:

- A measured volume of an acidic or basic solution of unknown concentration is placed in a beaker.(Initial pH of the solution is read and recorded)
- A buret is filled with the titrating solution of known concentration.This is called the standard solution, or titrant.
- Measured volumes of the standard solution are added slowly andmixed into the solution in the beaker. The pH is read and recordedafter each addition.
- This process continues until the reaction reaches the **equivalence point**.

The **equivalence point**:which is the point at which moles of H + ionfrom the acid equal moles of O H - ion from the base.

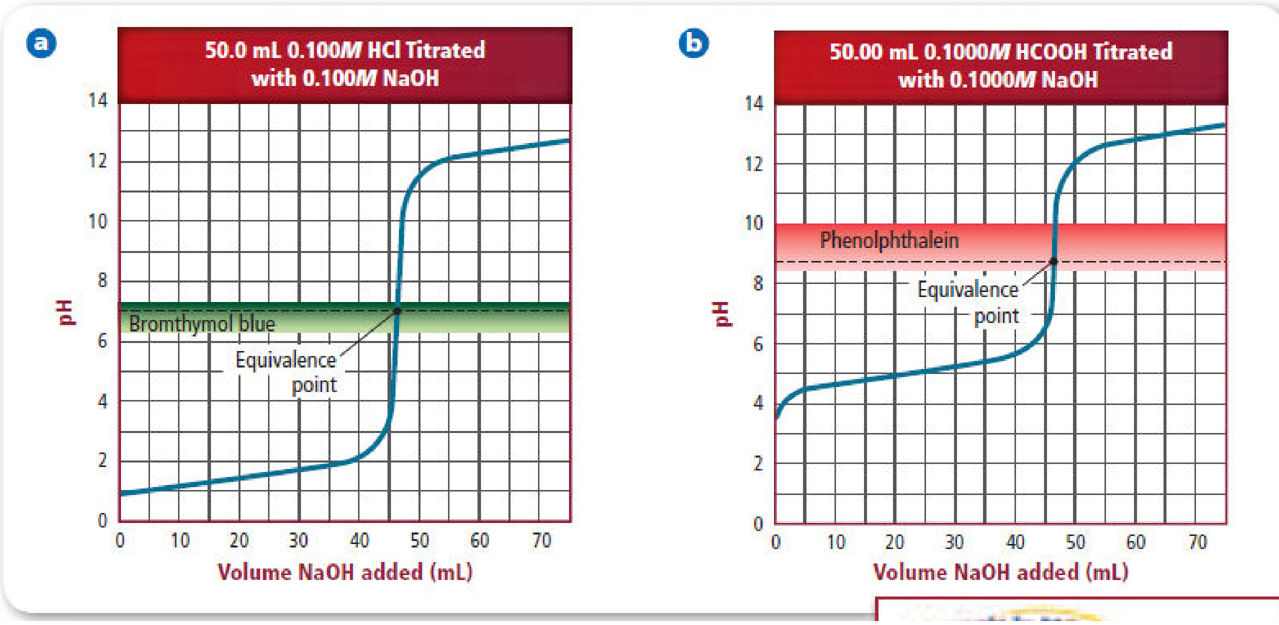


Figure 18.22a shows:

- How the pH of the solution changes during thetitration of 50.0 mL of 0.100M HCl, a strong acid, with 0.100M NaOH,a strong base.
- The initial pH of the 0.100M HCl is 1.00.
- As NaOHisadded, the acid is neutralized and the solution's pH increases gradually.
- However, when nearly all of the H + ions from the acid have been usedup, the pH increases dramatically with the addition of an exceedinglysmall volume of NaOH.
- This abrupt increase in pH occurs at the equivalencepoint of the titration.
- Beyond the equivalence point, the additionof more NaOH again results in a gradual increase in pH.
- **<u>Question</u>** :Identify two ways in which the graphs in Figure 18.22 are different.

Acid-base indicators:

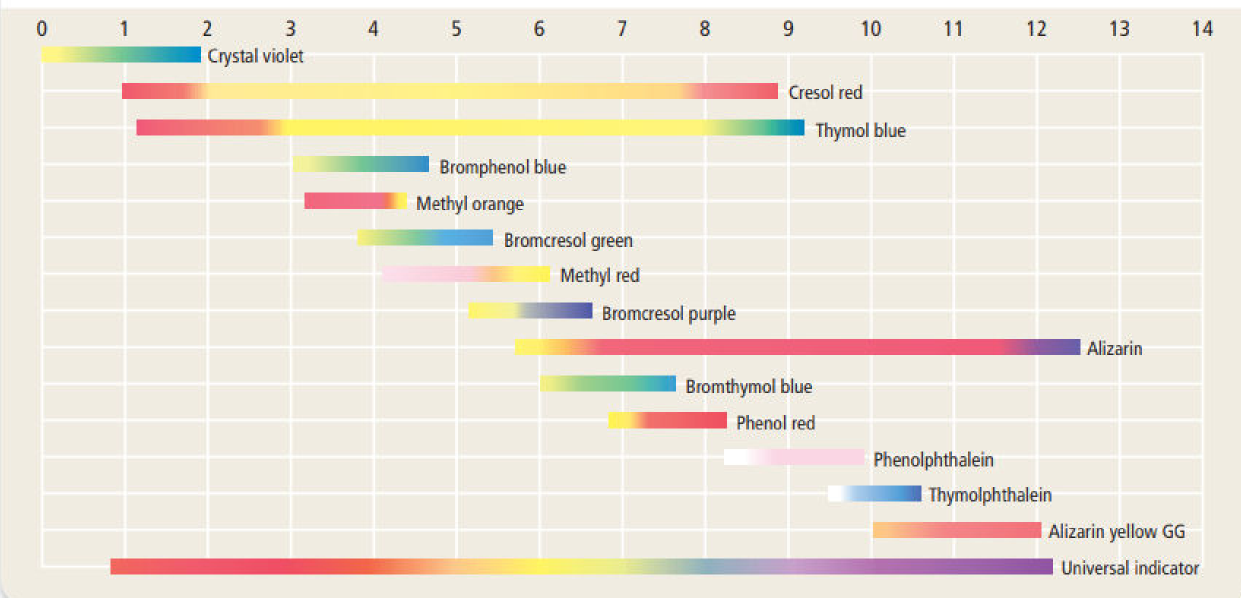
- Acid-base indicators: Chemical dyes whose colors are affected by acidic and basic solutions.
- Many natural substances act asindicators like tea.
- Tea contains compounds called polyphenolsthat have slightly ionizable hydrogen atoms and therefore are weakacids:

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

indicator molecules indicator ions

- Adding acid in the form of lemon juice to a cup of tea depressesthe ionization according to Le Châtelier's principle, and the color of theun-ionized polyphenols becomes more apparent.

Figure 18.24 Choosing the right indicator is important. The indicator must change color at the equivalence point of the titration which is not always at pH 7.



- As shown in Figure 18.22, bromthymol blue is a good choice for a titration of a strong acid with a strong base, and that phenophthalein changes color at the equivalence point of a titration of a weak acid with a strong base .

Indicators and titration end point:

- Many indicators used for titration are weak acids.
- Each has it own particular pH or pH ranges over which it changes color.
- **End point:** The point at which the indicator used in a titration changes color**.**
- It is important to choose an indicator for a titration that will change color at the equivalence point of the titration.

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.

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

1. 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).
$(M_B)\ (V_B) = (\text{mol/L})(\text{L}) = \text{mol NaOH in standard}$

2. 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.
$\text{mol } H_2SO_4 \text{ titrated} = \text{mol NaOH in standard} \times \dfrac{1 \text{ mol } H_2SO_4}{2 \text{ mol NaOH}}$

3. M_A represents the molarity of the acid and V_A represents the volume of the acid in liters. $M_A = \dfrac{\text{mol } H_2SO_4 \text{ titrated}}{V_A}$

Apply this strategy as you study Example Problem 18.6.

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44. 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 the acid solution?

45. What is the concentration of a household ammonia cleaning solution if 49.90 mL of 0.5900*M* HCl is required to neutralize 25.00 mL of the solution?

46. Challenge How many milliliters of 0.500*M* NaOH would neutralize 25.00 mL of 0.100*M* H_3PO_4?

95. How many milliliters of 0.225M HCl would be required to titrate 6.00 g of KOH? (KOH = 56 g/mol

Salt Hydrolysis:

- Salt Hydrolysis: means the reaction of salt ions with water that come from weak acids or bases.
- **Salts that produce basic solutions**: (hydrolysis of anion)
  For example, KF:
- The K^+ ions do not react with water, but the F^- ion is a weak BrønstedLowrybase.

$$F^-(aq) + H_2O(l) \rightleftharpoons HF(aq) + OH^-(aq)$$

- The production of the O H - ions makes the solution basic.
  - **Salts that produce acidic solutions**: (hydrolysis of cation)
- For example : NH_4Cl is the salt of a weak base (NH_3) and a strong acid (HCl).
- The Cl^- ions do not react with water, but the NH_4^+ ion is a weakBrønsted-Lowry acid. Ammonium ions react with water molecules:

$$NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$$

- The presence of hydronium ions makes the solution acidic.
  - **Salts that produce neutral solutions:**
- For example: Sodium nitrate ($NaNO_3$) is the salt of a strong acid (HNO_3) and a strong base (NaOH).
- No salt hydrolysis occurs because neither Na^+ nor NO_3^- react with water. Therefore, a solution of sodium nitrate is neutral.

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47. Write equations for the salt hydrolysis reactions occuring when the following salts dissolve in water. Classify each as acidic, basic, or neutral.

a. ammonium nitrate **c.** rubidium acetate
b. potassium sulfate **d.** calcium carbonate

48. Challenge Write the equation for the reaction that occurs in a titration of ammonium hydroxide (NH_4OH) with hydrogen bromide (HBr). Will the pH at the equivalence point be greater or less than 7?

47. a. $NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$
The solution is acidic.

b. $SO_4^{2-}(aq) + H_2O(l) \rightleftharpoons HSO_4^-(aq) + OH^-(aq)$
The solution is neutral.

c. $CH_3COO^-(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$
The solution is basic.

d. $CO_3^{2-}(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$
The solution is basic.

48. $NH_4OH(aq) + HBr(aq) \rightarrow NH_4Br(aq) + H_2O(l)$
$NH_4^+(aq) + H_2O(aq) \rightleftharpoons H_3O^+(aq) + NH_3$
Hydronium ions are formed so the pH will be less than 7.

Question : which ion react with water?

PO_4^{-3}	CO_3^{-2}	SO_4^{-2}	CH_3COO^-	K^+	NH_4^+	F^-	NO_3^-

- **H.W:** solve the questions 85 – 93 , 108 , 111

SECTION 4 REVIEW

49. Each neutralization reaction is the reaction of one mole of hydrogen ion with one mole of hydroxide to form one mole of water.

50. Equivalence point is the pH at which the moles of H^+ ions from the acid equal the moles of OH^- ions from the base. The end point is the point at which the indicator used in a titration changes color.

51. The pH of the unbuffered solution increases more than the pH of the buffered solution.

52. $M_A = 0.1214M$

53. Use ammonia and a salt of ammonia such as ammonium nitrate or ammonium chloride. Use equal molar amounts of the acid and its salt.

54. Place a measured volume of CsOH solution into a flask. Add an indicator such as bromothymol blue. Fill a buret with the 0.250M HNO_3 solution. Record the initial buret reading. Add HNO_3 solution slowly to the CsOH solution until the end point. Record the final buret reading. Calculate the volume of HNO_3 added. Use the volume and molarity of HNO_3 and the volume of CsOH to calculate the molarity of the CsOH solution. Refer to the Solutions Manual for the ionic equations.