شكراً لتحميلك هذا الملف من موقع المناهج الإماراتية

أسئلة مراجعة الوحدة الرابعة Bases and Acids الأحماض والقواعد

موقع المناهج ⇦ المناهج الإماراتية ⇦ الصف الثاني عشر المتقدم ⇦ كيمياء ⇦ الفصل الثاني ⇦ الملف

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التواصل الاجتماعي بحسب الصف الثاني عشر المتقدم

روابط مواد الصف الثاني عشر المتقدم على تلغرام

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12 Advanced Chemistry 2023/24

English Version

Chapter 4 Acids and Bases

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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

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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.

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

  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.

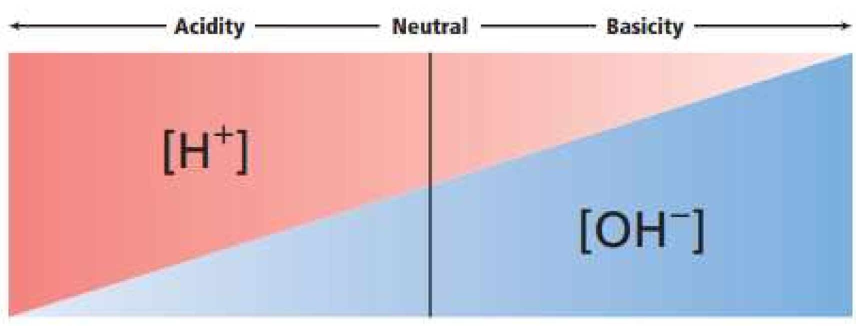


Figure 18.3 Note how [H⁺] and [OH⁻] change simultaneously. As [H⁺] decreases to the right, [OH⁻] increases to the right.
Identify *the point in the diagram at which the two ion concentrations are equal.*

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:

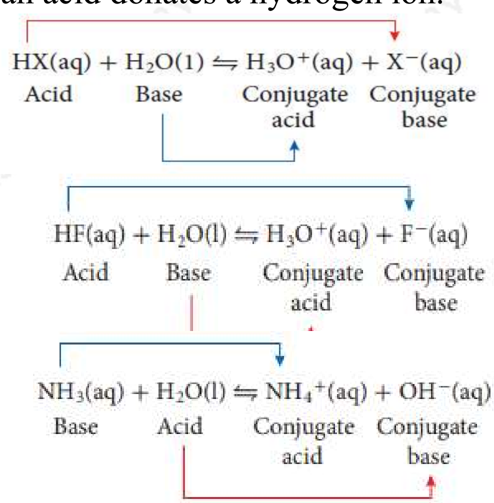
- An acid is a hydrogen-ion donor.
- A base is a hydrogenion acceptor.

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^- .**
- **Explain** how the ion $HCO3^-$ can be both an acid and a base (**amphoteric substance**).

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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 :

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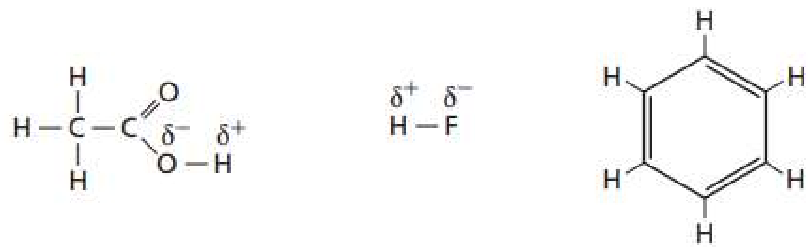
- A triprotic acid: An acid with three hydrogen ions to donate like H_3PO_4 , H_3BO_3.
  - Write the ionizations steps of phosphoric acid:

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.



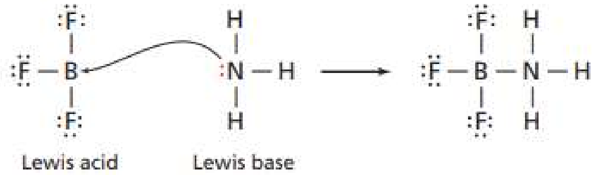
Acetic acid   Hydrogen fluoride   Benzene

Table 18.1 Some Common Acids and Their Conjugate Bases

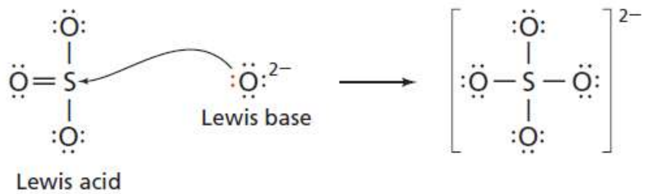
Acid		Congugate Base	
Name	Formula	Name	Formula
Hydrochloric acid	HCl	Chloride ion	Cl^-
Nitric acid	HNO_3	Nitrate ion	NO_3^-
Sulfuric acid	H_2SO_4	Hydrogen sulfate ion	HSO_4^-
Hydrogen sulfate ion	HSO_4^-	Sulfate ion	SO_4^{2-}
Hydrofluoric acid	HF	Fluoride ion	F^-
Hydrocyanic acid	HCN	Cyanide	CN^-
Acetic 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)

1- 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^+ .
2- 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 manyuses, 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.

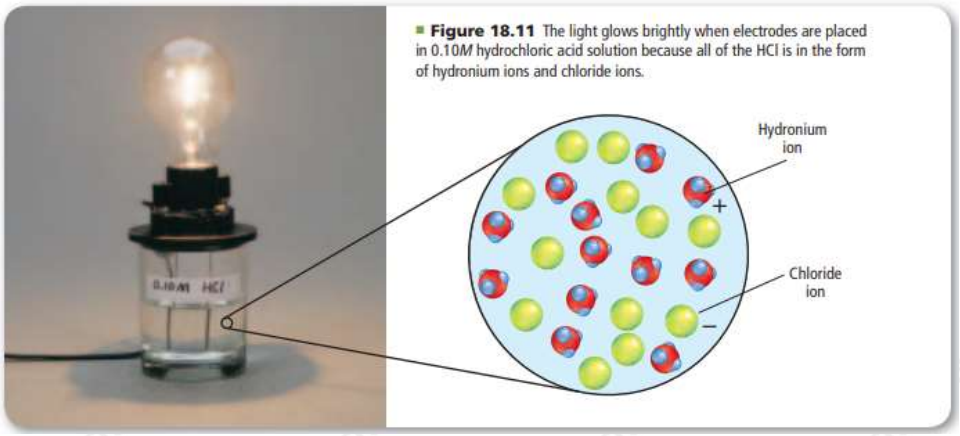
According to the **(14)** ____________________ model, an acid donates **(15)** ____________________ ions, and a base accepts **(16)** ____________________ ions. According to this model, in an acid-base reaction, each acid has a **(17)** ____________________, and each base has a **(18)** ____________________.

Section 2 : Strengths of of Acids and Bases:

<u>Strength of Acids:</u>

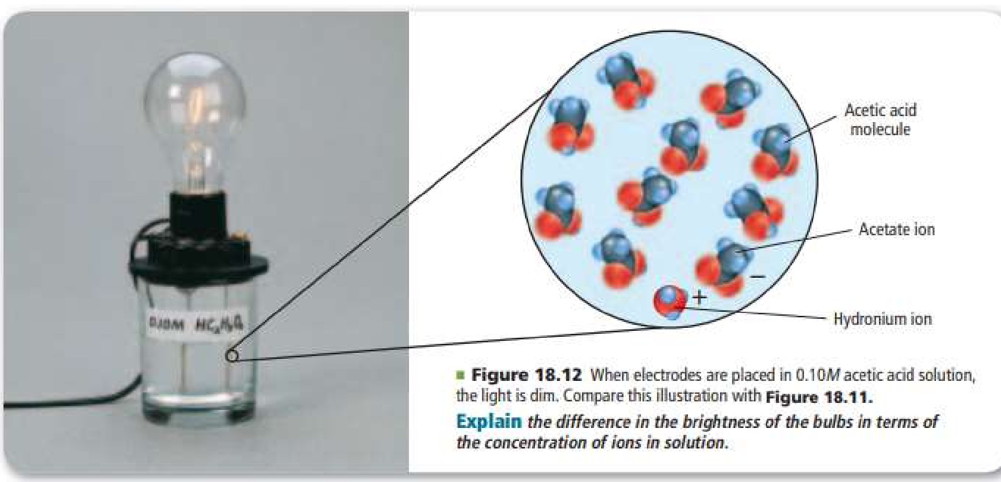
- **Degree of ionization:**
- Strong Acids: completely ionized (strong electrolytes) and good conductors , because strong acids produce the maximum number of ion.

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



■ **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 **left** 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 expression provides the quantitative measure of acid strength**.**
- The equilibrium constant, K_{eq} , provides a quantitative measure of the degree of ionization of the acid.
- Consider hydrocyanic acid (HCN), also known as **prussic acid** 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

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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 andhydroxide 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.

$$\underset{\text{Acid}}{CH_3NH_2(aq)} + \underset{\text{Base}}{H_2O(l)} \rightarrow \underset{\text{Conjugate acid}}{CH_3NH_3^+(aq)} + \underset{\text{Conjugate base}}{OH^-(aq)}$$

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

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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 (⟶) 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

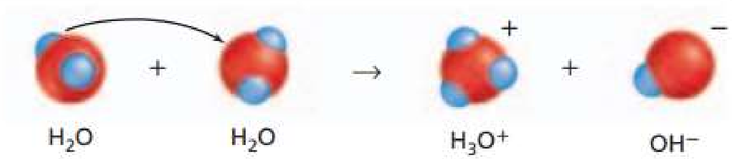
$$K_w = [H^+][OH^-]$$

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'sprinciple)

……

……

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.

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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.

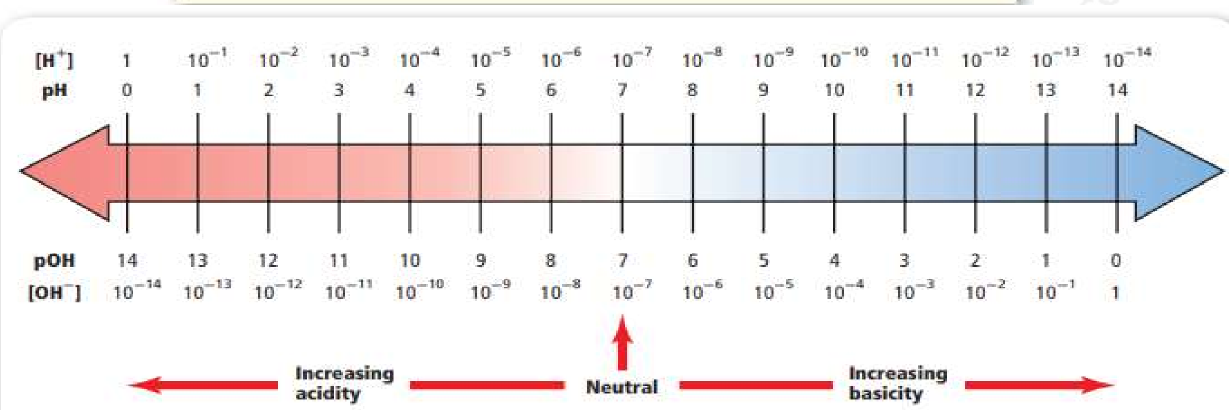

Figure 18.15 Study this diagram to sharpen your understanding of pH and pOH. Note that at each vertical position, the sum of pH (above the arrow) and pOH (below the arrow) equals 14. Also note that at every position, the product of $[H^+]$ and $[OH^-]$ equals 10^{-14}.

PRACTICE Problems

Extra Practice Page 989 and glencoe.com

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

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

PRACTICE Problems

Extra Practice Page 989 and glencoe.com

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. $[OH^-] = 0.000033M$

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

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 .

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

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.

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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.

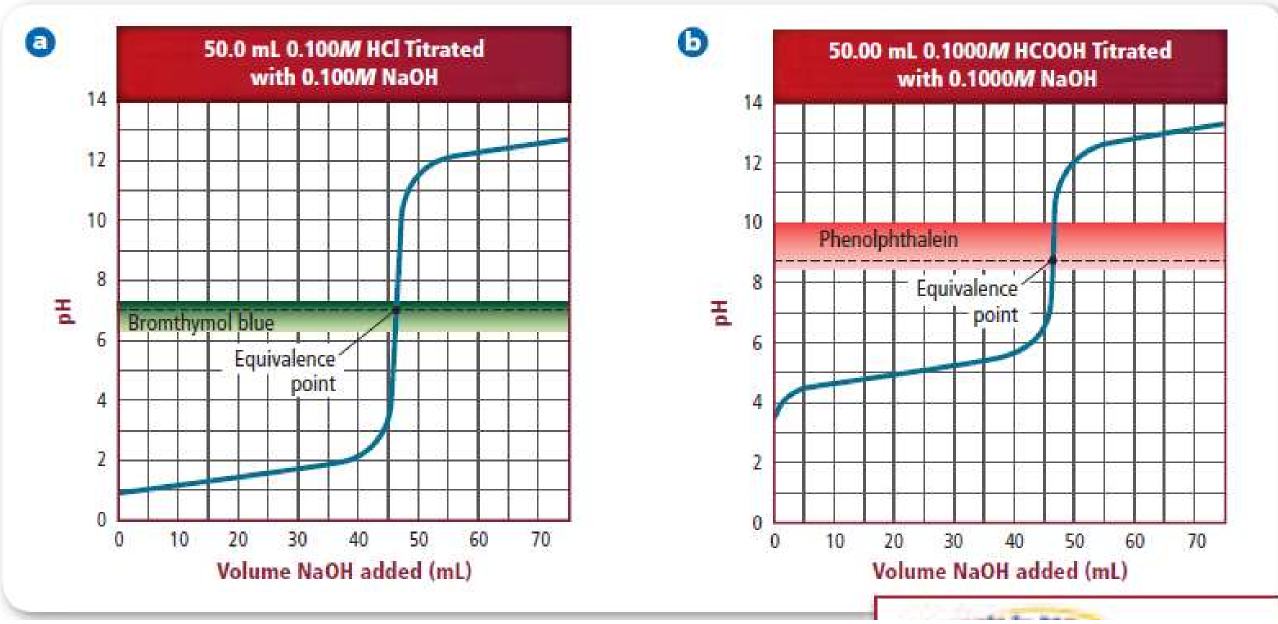


Figure 18.22a shows:

- 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:

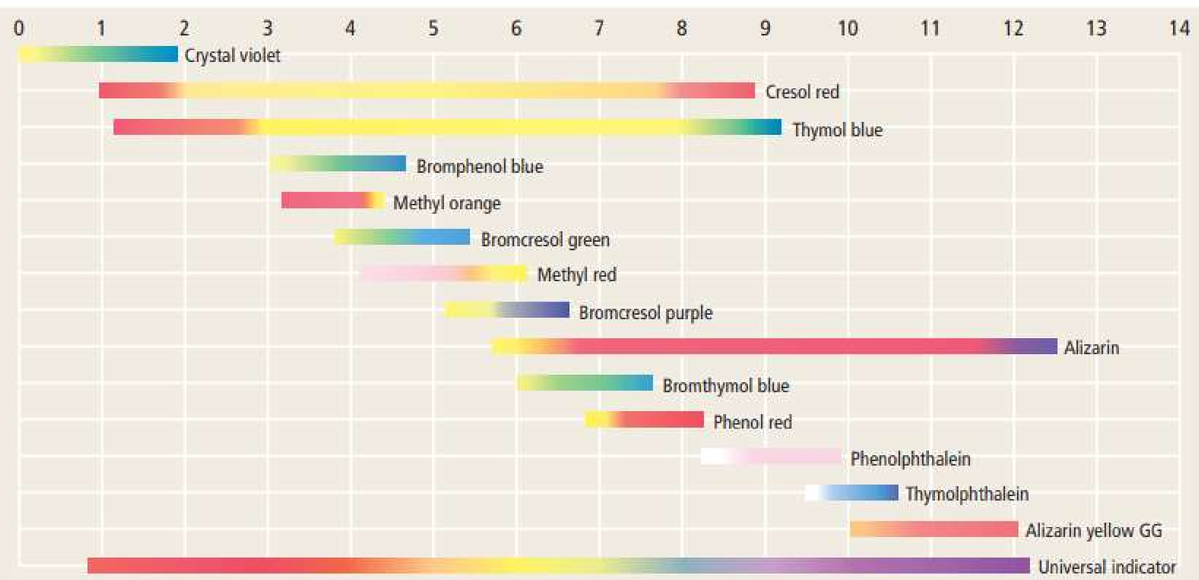
- Acid-base indicators: Chemical dyes whose colors are affected by acidic and basic solutions.
- Many natural substances act asindicators like tea.
- Tea contains compounds called polyphenolsthat have slightly ionizable hydrogen atoms and therefore are weakacids:

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

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 .

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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.

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

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).
   $(M_B)\ (V_B) = (mol/L)(L) = mol\ NaOH\ in\ standard$
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.
   $mol\ H_2SO_4\ titrated = mol\ NaOH\ in\ standard \times \frac{1\ mol\ H_2SO_4}{2\ mol\ NaOH}$
3. M_A represents the molarity of the acid and V_A represents the volume of the acid in liters. $M_A = \frac{mol\ H_2SO_4\ titrated}{V_A}$

Apply this strategy as you study Example Problem 18.6.

PRACTICE Problems

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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.

PRACTICE Problems Extra Practice Pages 989–990 and glencoe.com

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.

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English Version

Chapter 5 Redox Reactions

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Done by Teacher Abdul Raheem Qadomi

Manasra Academy

CHAPTER 5

Redox Reactions

Section1: Oxidation and Reduction:

- Oxidation-reduction reactions are also known as redox reactions.
- Redox reaction: a reaction in which electrons are transferred from one substance to another.
- It is a include the following reactions:
  – formation of a compound from its elements $2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$

  – all combustion reactions $2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

  – single replacement reactions $2KBr_{(aq)} + Cl_{2(aq)} \rightarrow 2KCl_{(aq)} + Br_{2(aq)}$

Oxidation and Reduction:

- **Oxidation** :is defined as the loss of electrons from atoms of a substance
  $Na \rightarrow Na^{+} + e^{-}$
- **Reduction** is defined as the gain of electrons by the atoms of a substance.
  $Cl_2 + 2e^{-} \rightarrow 2Cl^{-}$

Oxidation number:

- is the number of electrons lost or gained by the atom when it forms an ion.
- When an atom or ion is reduced, the numerical value of its oxidation number decreases.
- When an atom or ion is oxidized, its oxidation number increases.
- For example : determine which substance is oxidized or reduced in the following equation: $2K_{(s)} + Cl_{2(g)} \rightarrow 2KCl_{(s)}$

Oxidation number: +3 Ionic charge: 3+

Oxidizing and Reducing Agents:

- **Oxidizing agent** :The substance that oxidizes another substance by accepting its electrons (the substance that is reduced in a redox reaction).
- **Reducing agent :**The substance that reduces another substance by losing its electrons (the substance that is oxidized in a redox reaction).
- For example : determine the oxidizing agent and reducing agent:

$$2K_{(s)} + Cl_{2(g)} \rightarrow 2KCl_{(s)}$$

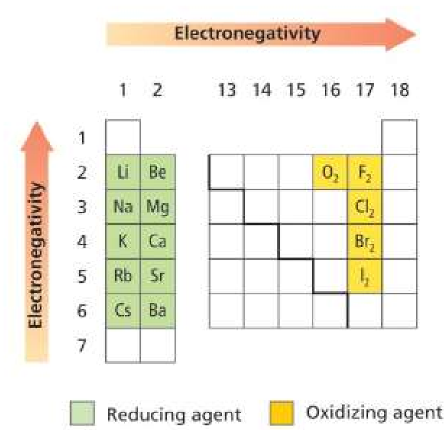
Redox and Electronegativity:

- Redox reactions are not limited to atoms of an element changing to ions.
- Some redox reactions involve changes in molecular substances.
- Which atom oxidized and which reduced in the following reaction:
  $N_{2(g)} + 3H_{2(g)} \rightarrow NH_{3(g)}$

- To determine which was oxidized and which was reduced, you must know which atom is more electronegative.
- Elements with high electronegativity, parial gain of electron, and so they are strong oxidizing agents.

Figure 4 The electronegativity of elements increases from left to right across the periodic table, and it decreases going down a group. Elements with low electronegativity are strong reducing agents, and elements with high electronegativity are strong oxidizing agents.

Predict ***which element would be the strongest oxidizing agent. Which is the strongest reducing agent?***



PRACTICE Problems

1. Identify each of the following changes as either oxidation or reduction. Recall that e^- is the symbol for an electron.
   a. $I_2 + 2e^- \rightarrow 2I^-$
   b. $K \rightarrow K^+ + e^-$
   c. $Fe^{2+} \rightarrow Fe^{3+} + e^-$
   d. $Ag^+ + e^- \rightarrow Ag$
2. Identify what is oxidized and what is reduced in the following processes.
   a. $2Br^- + Cl_2 \rightarrow Br_2 + 2Cl^-$
   b. $2Ce + 3Cu^{2+} \rightarrow 3Cu + 2Ce^{3+}$
   c. $2Zn + O_2 \rightarrow 2ZnO$
   d. $2Na + 2H^+ \rightarrow 2Na^+ + H_2$
3. Identify the oxidizing agent and the reducing agent in the following equation. Explain your answer.
   $$Fe(s) + 2Ag^+(aq) \rightarrow Fe^{2+}(aq) + 2Ag(s)$$
4. **Challenge** Identify the oxidizing agent and the reducing agent in each reaction.
   a. $Mg + I_2 \rightarrow MgI_2$
   b. $H_2S + Cl_2 \rightarrow S + 2HCl$

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Determining Oxidation Numbers:

- To understand all types of redox reactions, the oxidation number of the atoms involved in the reaction must be determined.
- The following table outlines the rules chemists use to make this determination easier.

Table 19.2 Rules for Determining Oxidation Numbers

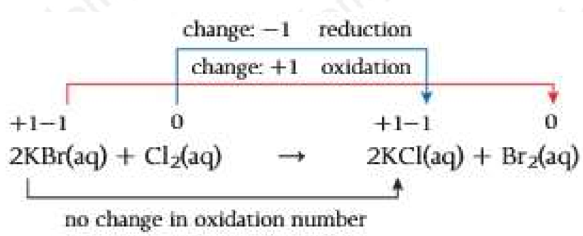
Rule	Example	$n_{element}$
1. The oxidation number of an uncombined atom is zero.	Na, O_2, Cl_2, H_2	0
2. The oxidation number of a monatomic ion is equal to the charge of the ion.	Ca^{2+}	+2
	Br^-	−1
3. The oxidation number of the more-electronegative atom in a molecule or a complex ion is the same as the charge it would have if it were an ion.	N in NH_3	−3
	O in NO	−2
4. The oxidation number of the most-electronegative element, fluorine, is always −1 when it is bonded to another element.	F in LiF	−1
5. The oxidation number of oxygen in compounds is always −2 except in peroxides, such as hydrogen peroxide (H_2O_2), where it is −1. When it is bonded to fluorine, the only element more electro-negative than oxygen, the oxidation number of oxygen is positive.	O in NO_2	−2
	O in H_2O_2	−1
6. The oxidation number of hydrogen in most of its compounds is +1, except in metal hydrides; then, the oxidation number is −1.	H in NaH	−1
7. The oxidation numbers of group 1 and 2 metals and aluminum are positive and equal to their number of valence electrons.	K	+1
	Ca	+2
	Al	+3
8. The sum of the oxidation numbers in a neutral compound is zero.	$CaBr_2$	(+2) + 2(−1) = 0
9. The sum of the oxidation numbers of the atoms in a polyatomic ion is equal to the charge of the ion.	SO_3^{2-}	(+4) + 3(−2) = −2

PRACTICE Problems

5. Determine the oxidation number of the boldface element in the following formulas for compounds.
   a. $NaClO_4$ b. $AlPO_4$ c. HNO_2
6. Determine the oxidation number of the boldface element in the following formulas for ions.
   a. NH_4^+ b. AsO_4^{3-} c. CrO_4^{2-}
7. Determine the oxidation number of nitrogen in each of these molecules.
   a. NH_3 b. KCN c. N_2H_4
8. **Challenge** Determine the net change of oxidation number of each of the elements in these redox equations.
   a. $C + O_2 \rightarrow CO_2$
   b. $Cl_2 + ZnI_2 \rightarrow ZnCl_2 + I_2$
   c. $CdO + CO \rightarrow Cd + CO_2$

5. a. +7
   b. +5
   c. +3
6. a. −3
   b. +5
   c. +6
7. a. −3
   b. −3
   c. −2
8. a. C, +4; O, −2
   b. I, +1; Cl, −1; Zn, no change
   c. C, +2; Cd, −2; O, no change

Oxidation Numbers in Redox Reactions:



- Oxidation-reduction reactions are changes in oxidation number.
- Atoms that are reduced have their oxidation number decreased.
- Atoms that are oxidized have their oxidation number increased.

H.W : solve the questions 33 – 50 .

SECTION 1 REVIEW

9. If an atom or ion loses an electron, some other species must gain the electron.
10. An oxidizing agent causes another species to be oxidized by gaining the electrons from it. A reducing agent causes another species to be reduced by losing electrons to that element.
11. $2Fe(s) + 6HBr(aq) \rightarrow 2FeBr_3(aq) + 3H_2(g)$; Fe is oxidized, H is reduced
12. a. +5
    b. −3
    c. +5
    d. +6
13. a. +7
    b. +7
    c. +3
    d. −3
14. In general, as you move down the periodic table within a family, the tendency to lose electrons increases, so the reducing ability increases.

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

1. Redox reactions are characterized by
- **a.** formation of a solid, a gas, or water.
- **b.** replacement of one element in a compound by another element.
- **c.** sharing of electrons.
- **d.** transfer of electrons.

2. If a calcium atom loses two electrons, it becomes
- **a.** a Ca^{2-} ion.
- **b.** an oxidizing agent.
- **c.** oxidized.
- **d.** reduced.

3. In a redox reaction, an oxidizing agent is
- **a.** balanced.
- **b.** increased in oxidation number.
- **c.** oxidized.
- **d.** reduced.

4. An oxidation reaction occurs
- **a.** at the same time a reduction reaction occurs.
- **b.** before its corresponding reduction reaction occurs.
- **c.** independently of any reduction reaction.
- **d.** only when electrons are gained.

5. Consider the equation $Ca(s) + O_2(g) \rightarrow 2CaO(s)$.
In this reaction, calcium is oxidized because it
- **a.** becomes part of a compound.
- **b.** gains electrons.
- **c.** loses electrons.
- **d.** reacts with oxygen.

6. The number of electrons lost by an element when it forms ions is the element's
- **a.** charge.
- **b.** oxidation number.
- **c.** reduction number.
- **d.** shared electrons.

7. A loss of electrons is

a. oxidation. **b.** oxidation–reduction. **c.** redox. **d.** reduction.

8. Redox reactions can involve

a. ions only.
b. molecules only.
c. uncharged atoms only.
d. ions, molecules, or uncharged atoms.

Use your answers from questions 9–15 to fill in the following table for the listed reactions. For each reaction, show what is oxidized, what is reduced, the oxidizing agent, and the reducing agent.

Equation	Oxidized	Reduced	Oxidizing Agent	Reducing Agent
16. $Cd(s) + NiO(s) \rightarrow CdO(s) + Ni(s)$				
17. $Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$				
18. $2Sb(s) + 3I_2(g) \rightarrow 2SbI_3(s)$				
19. $2Cu_2S(s) + 3O_2(g) \rightarrow 2Cu_2O(s) + 2SO_2(g)$				
20. $PbO_2(s) + Pb(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(aq) + 2H_2O(l)$				
21. $NH_4NO_3(s) \rightarrow 2H_2O(g) + N_2O(g)$				
22. $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$				

Section 2 :

Balancing Redox Equations Using Half-Reaction:

- In chemistry, a **species** is any kind of chemical unit involved in a process.
- Oxidation-reduction reactions occur whenever a species that can give up electrons comes in contact with another species that can accept them.

Table 6 The Half-Reaction Method

1. Write the unbalanced, net ionic equation for the reaction, omitting spectator ions. $Fe + Cu^{2+} + SO_4^{2-} \rightarrow Cu + 2Fe^{3+} + 3SO_4^{2-}$ $Fe + Cu^{2+} \rightarrow Cu + 2Fe^{3+}$
2. Write separate, incomplete equations for the oxidation and reduction half-reactions, including oxidation numbers. $\overset{0}{Fe} \rightarrow \overset{+3}{2Fe^{3+}}$ (oxidation) $\overset{+2}{Cu^{2+}} \rightarrow \overset{0}{Cu}$ (reduction)
3. Balance the atoms in the half-reactions. Balance the charges in each half-reaction by adding electrons as reactants or products. $2Fe \rightarrow 2Fe^{3+} + 6e^-$ $Cu^{2+} + 2e^- \rightarrow Cu$
4. Adjust the coefficients so that the number of electrons lost in oxidation equals the number of electrons gained in reduction. $2Fe \rightarrow 2Fe^{3+} + 6e^-$ $3Cu^{2+} + 6e^- \rightarrow 3Cu$
5. Add the half-reactions and cancel or reduce like terms on both sides of the equation. $2Fe + 3Cu^{2+} \rightarrow 3Cu + 2Fe^{3+}$
6. Return spectator ions, if desired. Restore state descriptions. $2Fe(s) + 3CuSO_4(aq) \rightarrow 3Cu(s) + Fe_2(SO_4)_3(aq)$

PRACTICE Problems

Use the half-reaction method to balance the redox equations. Begin by writing the oxidation and reduction half-reactions. Leave the balanced equation in ionic form.

23. $Cr_2O_7^{2-}(aq) + I^-(aq) \rightarrow Cr^{3+}(aq) + I_2(s)$ (in acidic solution)

24. $Mn^{2+}(aq) + BiO_3^-(aq) \rightarrow MnO_4^-(aq) + Bi^{2+}(aq)$ (in acidic solution)

25. Challenge $N_2O(g) + ClO^-(aq) \rightarrow NO_2^-(aq) + Cl^-(aq)$ (in basic solution)

23. $14H^+(aq) + Cr_2O_7^{2-}(aq) + 6I^- \longrightarrow 3I_2(s) + 2Cr^{3+}(aq) + 7H_2O(l)$

24. $3Mn^{2+}(aq) + 5BiO_3^-(aq) + 6H^+(aq) \longrightarrow 3MnO_4^-(aq) + 5Bi^{2+}(aq) + 3H_2O(l)$

25. $N_2O(g) + 2ClO^-(aq) + 2OH^-(aq) \longrightarrow 2NO_2^-(aq) + 2Cl^-(aq) + H_2O(l)$

H.W : Solve questions 51 – 71 .

SECTION 2 REVIEW

26. Because the nucleus (specifically, number of protons) never changes during this type of reaction, whenever there is a transfer of electrons to or from a chemical species, there is a change in the net charge of that species. Oxidation increases the oxidation number, reduction reduces it.

27. It is important to know that H_2O and either H^+ or OH^- are available to balance the equation.

28. Answers should be similar to the information in Table 4.

29. An oxidation half reaction shows how many electrons a species loses. A reduction half reaction shows how many electrons a species gains.

30. oxidation: $Pb \rightarrow Pb^{2-} + 2e^-$
reduction: $Pd^{2+} + 2e^- \rightarrow Pd$

31. three Sn^{2+} ions; two Au^{3+} ions

32. **a.** $3HClO_3 \rightarrow 2ClO_2 + HClO_4 + H_2O$
b. $5H_2SeO_3 + 2HClO_3 \rightarrow 5H_2SeO_4 + Cl_2 + H_2O$
c. $Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \rightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$

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

______________ **1.** A *species* is any kind of chemical unit involved in a process.

______________ **2.** Glucose and sucrose are different types of sugars. A solution of glucose, sucrose, and water contains exactly *two* different species.

______________ **3.** A half-reaction is part of a *decomposition* reaction.

______________ **4.** When magnesium reacts with oxygen, $Mg \rightarrow Mg^{2+} + 2e^-$ is the *reduction* half of the reaction.

______________ **5.** A species that undergoes *oxidation* will donate electrons to any atom that accepts them.

______________ **6.** A species can be a molecule, an atom, or an *electron*.

______________ **7.** Balancing equations by half-reaction is based on the number of *atoms* transferred.

______________ **8.** Balancing half-reactions involves balancing *both atoms and charge*.

______________ **9.** In writing an equation in ionic form, ionic compounds are written as *molecules*.

______________ **10.** The half-reaction $SO_2 + H_2O + 2e^- \rightarrow SO_4^{2-} + 4H^+$ shows that the reaction takes place in *a basic* solution.

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English Version

Chapter 6 Electrochemistry

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Done by Teacher Abdul Raheem Qadomi

Manasra Academy

CHAPTER 6 / Electrochemistry

Section 1 : Votaic Cells

Redox in electrochemistry

- Electrochemistry : is the study of the redox processes by which chemical energy is converted to electrical energy and vise versa.
- study the following redox reaction and ansure the questions:

$$\overset{2e^-}{\overbrace{Zn(s) + Cu^{2+}}}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

- Write the oxidation half reaction: --

- Write the reduction half reaction : ---

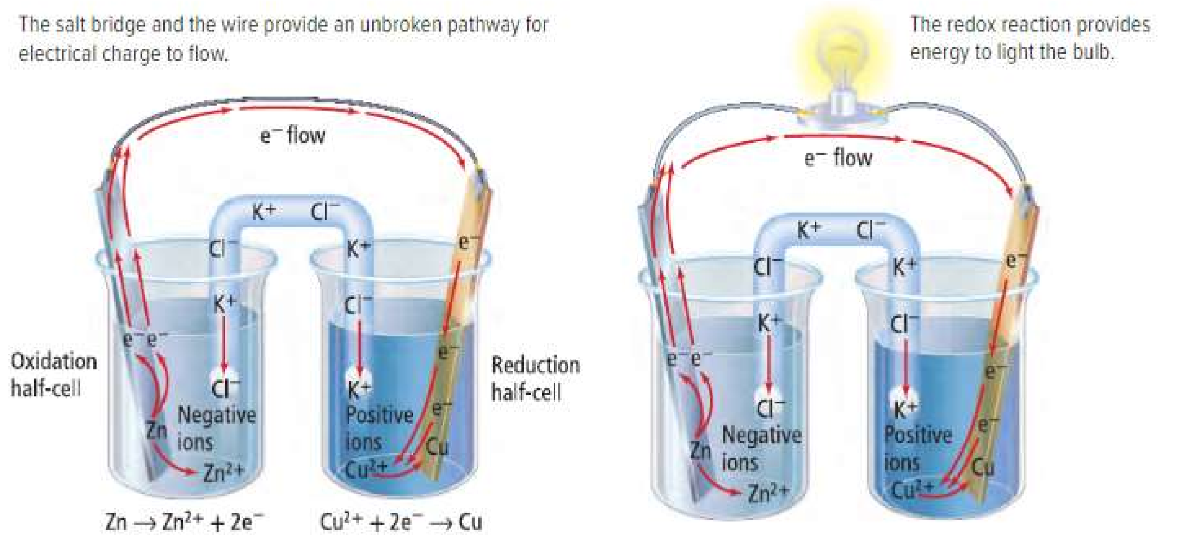
- What happened when we put a strip of zinc in $CuSO_4$ solution ?

- What is the type of energy which is produced from this reaction ?

- How can you generate electrical energy instead of heat energy ?

- What is a salt bridge? --

Electrochemical cells:



■ **Figure 2** The addition of a salt bridge completes the pathway. Negative ions move through the salt bridge to the zinc side. Positive ions move through the bridge to the copper side.

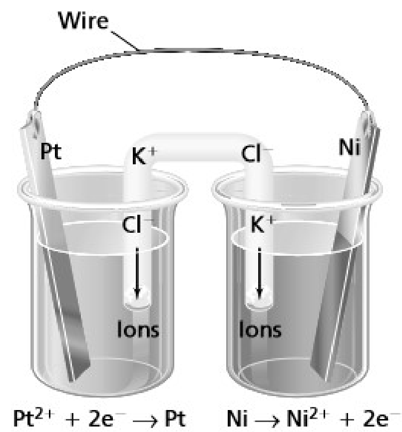
- **Electrochemical cell :** is an apparatus that uses a redox reaction to produce electrical energy or uses electrical energy to cause a chemical reaction **.**
- **Voltaic cell :** ……………………………………………………………………………………………
………

Chemistry of voltaic cells :

- An electrochemical cell consists of tow parts called ………………………………………………..
- Each half – cell contains ……………………………….. and ……………………………………………..
- **An electrode**: is an ……………………………………………………. , usually a metallic strip or graphite.
- **The anode :** ………………………………………………………………………………………………
- **The cathode** : ……………………………………………………………………………………………

voltaic	electrochemical cell	electric current	salt bridge	galvanic

Oxidation and reduction reactions can occur in separate solutions, as long as there are two connections between the solutions. One connection is a(n) **(1)**____________________ through which ions can flow. The other connection is a metal wire through which electrons can flow. The flow of ions or electrons is known as a(n) **(2)**____________________. The complete setup, called a(n) **(3)** ____________________, can convert chemical energy into electrical energy or electrical energy into chemical energy. These cells are also known as **(4)**____________________ cells or **(5)**____________________ cells.



Use the diagram of an electrochemical cell to answer the following questions.

6. The equation at the bottom of each beaker shows the half-reaction that is occurring in that beaker. What kind of reaction (oxidation or reduction) is occurring in each beaker?

Left beaker ____________________

Right beaker ____________________

7. Write the net ionic equation for this electrochemical cell.

____

8. In which direction do electrons move through the wire?

____

9. What kind of ions (positive or negative) move from the ∩-shaped tube into each beaker?

Left beaker ____________________ Right beaker ____________________

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Voltaic cells and energy:

- Electric charge can flow between tow points when a difference in electric potential energy exists between the tow points.
- Electromotive force (EMF): is the force that pushes electrons generated at the anode toward the cathode.
- A volt : is a unit used to measure cell potential.
- The voltage of a cell is determined by comparing the difference in the tendency of the tow electrode materials to accept electrons .
- The greater the difference , the greater the potential energy difference between the tow electrodes and the larger the voltage of the cell will be.

Calculating electrochemical cell potential :

- **Reduction potential:**the tendency of a substance to gain electrons .
- The reduction potential of an electrode cannot be determined directly because the reduction half – reaction must be coupled with an oxidation half – reaction.

The standered hydrogen electrode:

- It consists of :

  ………………………………….
  ………………………………….
  ………………………………….
  ………………………………….
  ………………………………….
  ………………………………….

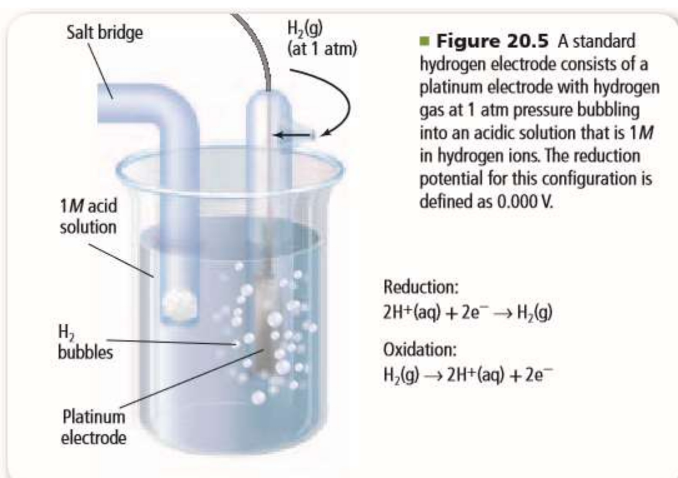

Figure 20.5 A standard hydrogen electrode consists of a platinum electrode with hydrogen gas at 1 atm pressure bubbling into an acidic solution that is 1*M* in hydrogen ions. The reduction potential for this configuration is defined as 0.000 V.

- The potetial also called , the standered reduction potential (E°) , of this standered hydrogen electrode is defined as 0.00 V .

$$2H^+(aq) + 2e^- \underset{\text{Oxidation}}{\overset{\text{Reduction}}{\rightleftharpoons}} H_2(g) \qquad E^0 = 0.000\ V$$

Half – cell potential:

- Over the years chemists have measured and recorded the standered reduction potetials of many different half – cells .
- The following table lists some common half – cell reactions in order of increasing reduction potential.
- The values in the table were obtained by measuring the potential when each half – cell was connected to a standered hydrogen half – cell .
- All of the half – reactions are written as reductions .
- The half – reaction that is more positive will proceed as a reduction , and the half – reaction that is more negative will proceed as an oxidation .
- Standered conditions : 1M solution of its ions , at 25°C and 1 atm.

Table 20.1 Standard Reduction Potentials

Half-Reaction	E^0 (V)	Half-Reaction	E^0 (V)
$Li^+ + e^- \rightarrow Li$	−3.0401	$Cu^{2+} + e^- \rightarrow Cu^+$	+0.153
$Ca^{2+} + 2e^- \rightarrow Ca$	−2.868	$Cu^{2+} + 2e^- \rightarrow Cu$	+0.3419
$Na^+ + e^- \rightarrow Na$	−2.71	$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$	+0.401
$Mg^{2+} + 2e^- \rightarrow Mg$	−2.372	$I_2 + 2e^- \rightarrow 2I^-$	+0.5355
$Be^{2+} + 2e^- \rightarrow Be$	−1.847	$Fe^{3+} + e^- \rightarrow Fe^{2+}$	+0.771
$Al^{3+} + 3e^- \rightarrow Al$	−1.662	$NO_3^- + 2H^+ + e^- \rightarrow NO_2 + H_2O$	+0.775
$Mn^{2+} + 2e^- \rightarrow Mn$	−1.185	$Hg_2^{2+} + 2e^- \rightarrow 2Hg$	+0.7973
$Cr^{2+} + 2e^- \rightarrow Cr$	−0.913	$Ag^+ + e^- \rightarrow Ag$	+0.7996
$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$	−0.8277	$Hg^{2+} + 2e^- \rightarrow Hg$	+0.851
$Zn^{2+} + 2e^- \rightarrow Zn$	−0.7618	$2Hg^{2+} + 2e^- \rightarrow Hg_2^{2+}$	+0.920
$Cr^{3+} + 3e^- \rightarrow Cr$	−0.744	$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$	+0.957
$S + 2e^- \rightarrow S^{2-}$	−0.47627	$Br_2(l) + 2e^- \rightarrow 2Br^-$	+1.066
$Fe^{2+} + 2e^- \rightarrow Fe$	−0.447	$Pt^{2+} + 2e^- \rightarrow Pt$	+1.18
$Cd^{2+} + 2e^- \rightarrow Cd$	−0.4030	$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	+1.229
$PbI_2 + 2e^- \rightarrow Pb + 2I^-$	−0.365	$Cl_2 + 2e^- \rightarrow 2Cl^-$	+1.35827
$PbSO_4 + 2e^- \rightarrow Pb + SO_4^{2-}$	−0.3588	$Au^{3+} + 3e^- \rightarrow Au$	+1.498
$Co^{2+} + 2e^- \rightarrow Co$	−0.28	$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	+1.507
$Ni^{2+} + 2e^- \rightarrow Ni$	−0.257	$Au^+ + e^- \rightarrow Au$	+1.692
$Sn^{2+} + 2e^- \rightarrow Sn$	−0.1375	$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$	+1.776
$Pb^{2+} + 2e^- \rightarrow Pb$	−0.1262	$Co^{3+} + e^- \rightarrow Co^{2+}$	+1.92
$Fe^{3+} + 3e^- \rightarrow Fe$	−0.037	$S_2O_8^{2-} + 2e^- \rightarrow 2SO_4^{2-}$	+2.010
$2H^+ + 2e^- \rightarrow H_2$	**0.0000**	$F_2 + 2e^- \rightarrow 2F^-$	+2.866

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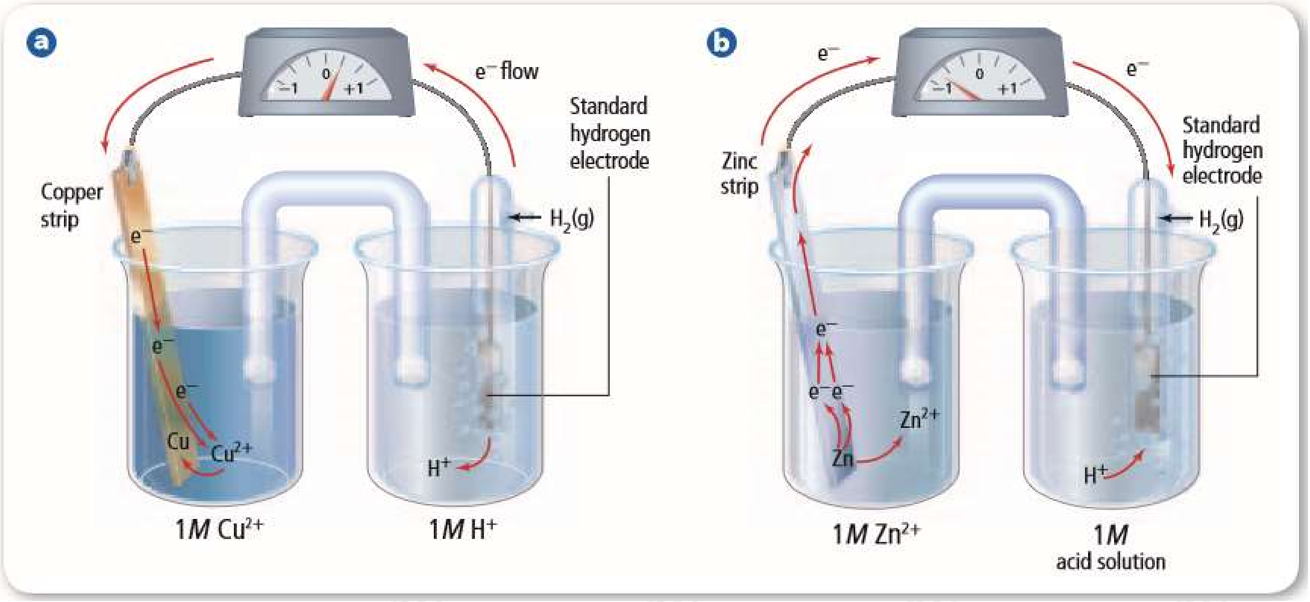


Figure 20.6 a. When a Cu | Cu^{2+} electrode is connected to the hydrogen electrode, electrons flow toward the copper strip and reduce Cu^{2+} ions to Cu atoms. The voltage of this reaction is +0.342 V. **b.** When a Zn | Zn^{2+} electrode is connected to the hydrogen electrode, electrons flow away from the zinc strip and zinc atoms are oxidized to Zn^{2+} ions. The voltage of this reaction is −0.762 V.

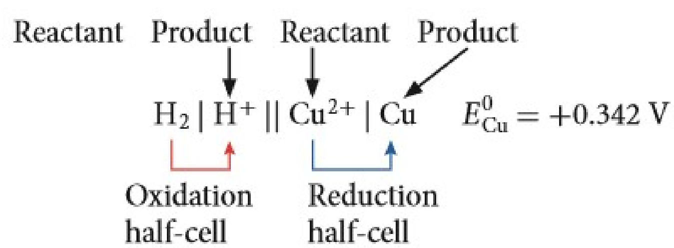
Determining electrochemical cell potentials:

- To calculate the electric potential of a voltic cell ($Zn/Zn^{+2}//Cu^{+2}/Cu$) :
- The first step is to determine (E°_{Cu}) :

$$H_2(g) \rightarrow 2H^+(aq) + 2e^- \text{ (oxidation half-cell reaction)}$$
$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s) \text{ (reduction half-cell reaction)}$$

$$H_2(g) + Cu^{2+}(aq) \rightarrow 2H^+(aq) + Cu(s) \text{ (overall redox reaction)}$$

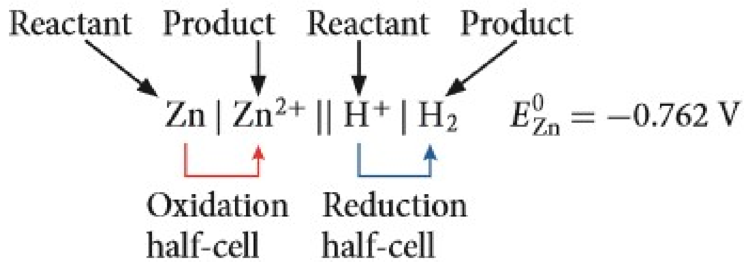
This reaction can be written in a form called cell notation.



- The next step is to determine (E°_{Zn}):

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^- \text{ (oxidation half-cell reaction)}$$
$$2H^+(aq) + 2e^- \rightarrow H_2(g) \text{ (reduction half-cell reaction)}$$
$$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g) \text{ (overall redox cell reaction)}$$

This reaction can be written in the following cell notation.



- The final step is to calculate E°_{cell} :

Formula for Cell Potential

$$E^0_{cell} = E^0_{reduction} - E^0_{oxidation}$$

E^0_{cell} represents the overall standard cell potential.
$E^0_{reduction}$ represents the standard half-cell potential for the reduction.
$E^0_{oxidation}$ represents the standard half-cell potential for the oxidation.

The standard potential of a cell is the standard potential of the reduction half-cell minus the standard potential of the oxidation half-cell.

$$E^0_{cell} = E^0_{Cu^{2+}|\,Cu} - E^0_{Zn^{2+}|\,Zn}$$
$$= +0.342 \text{ V} - (-0.762 \text{ V})$$
$$= +1.104 \text{ V}$$

PRACTICE Problems

Extra Practice Page 991 and glencoe.com

For each of these pairs of half-reactions, write the balanced equation for the overall cell reaction, and calculate the standard cell potential. Describe the reaction using cell notation. Refer to Chapter 19 to review writing and balancing redox equations.

1. $Pt^{2+}(aq) + 2e^- \rightarrow Pt(s)$ and $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$
2. $Co^{2+}(aq) + 2e^- \rightarrow Co(s)$ and $Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$
3. $Hg^{2+}(aq) + 2e^- \rightarrow Hg(l)$ and $Cr^{2+}(aq) + 2e^- \rightarrow Cr(s)$
4. **Challenge** Write the balanced equation for the cell reaction and calculate the standard cell potential for the reaction that occurs when these half-cells are connected. Describe the reaction using cell notation.

   $NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$

   $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

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Using standard reduction potentials:

- To calculate the standard potential of a voltaic cells .
- To determine if a proposed reaction will be spontaneous , if the E°_{cell} is positive the reaction is spontaneous .

PRACTICE Problems Extra Practice Page 991 and glencoe.com

Calculate the cell potential to determine if each of the following balanced redox reactions is spontaneous as written. Use Table 20.1 to help you determine the correct half-reactions.

5. $Sn(s) + Cu^{2+}(aq) \rightarrow Sn^{2+}(aq) + Cu(s)$

6. $Mg(s) + Pb^{2+}(aq) \rightarrow Pb(s) + Mg^{2+}(aq)$

7. $2Mn^{2+}(aq) + 8H_2O(l) + 10Hg^{2+}(aq) \rightarrow 2MnO_4^{-}(aq) + 16H^{+}(aq) + 5Hg_2^{2+}(aq)$

8. $2SO_4^{2-}(aq) + Co^{2+}(aq) \rightarrow Co(s) + S_2O_8^{2-}(aq)$

9. Challenge Using **Table 20.1,** write the equation and determine the cell voltage (E^0) for the following cell. Is the reaction spontaneous?

$$Al \mid Al^{3+} \parallel Hg^{2+} \mid Hg_2^{2+}$$

SECTION 1 REVIEW

10. An electrochemical cell in which an oxidation half-reaction and a reduction half-reaction are connected by a salt bridge results in a flow of electrons (electric current) through a conducting wire.
11. A voltaic cell consists of an anode, a cathode, a salt bridge, and a connecting wire between the two electrodes. Oxidation takes place at the anode, reduction takes place at the cathode, the salt bridge allows movement of ions from one solution to the other, and the wire allows the passage of electrons from the anode to the cathode.
12. a. $2Ag^+ + Ni \rightarrow 2Ag + Ni^{2+}$
    b. $Mg + 2H^+ \rightarrow Mg^{2+} + H_2$
    c. $2Fe^{3+}(aq) + 3Sn(s) \rightarrow 2Fe(s) + 3Sn^{2+}(aq)$
    d. $Pb(s) + 2I^-(aq) + Pt^{2+}(aq) \rightarrow PbI_2(s) + Pt(s)$
13. a. $E^0_{cell} = -2.004$ V, nonspontaneous
    b. $E^0_{cell} = +0.698$ V, spontaneous
    c. $E^0_{cell} = +1.178$ V, spontaneous
14. Concept Maps will vary. Refer to the Solutions Manual.

H.W : Solve questions 30 – 42 page 264 .

For each item in Column A, write the letter of the matching item in Column B.

Column A	Column B
______ **10.** One of the two parts of an electrochemical cell, where either oxidation or reduction takes place	**a.** battery
______ **11.** An electrode where oxidation takes place	**b.** electrical potential
______ **12.** An electrode where reduction takes place	**c.** half-cell
______ **13.** One or more electrochemical cells in a single package that generates electrical current	**d.** cathode
______ **14.** A measure of the amount of current that can be generated from an electrochemical cell to do work	**e.** anode

In your textbook, read about calculating cell electrochemical potential.

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

15. The tendency of an electrode to gain electrons is called
- **a.** electron potential.
- **b.** gravitational potential.
- **c.** reduction potential.
- **d.** oxidation potential.

16. A sheet of platinum covered with finely divided platinum particles is immersed in a 1*M* HCl solution containing hydrogen gas at a pressure of 1 atm and a temperature of 25°C. The platinum sheet is known as a
- **a.** standard platinum electrode.
- **b.** standard hydrogen electrode.
- **c.** hydrogen chloride electrode.
- **d.** platinum chloride electrode.

17. The standard reduction potential of a half-cell is a measure of
- **a.** concentration.
- **b.** pressure.
- **c.** temperature.
- **d.** voltage.

18. Which of the following is the correct way to represent the equation, $H_2(g) + Cu^{2+}(aq) \longrightarrow 2H^+(aq) + Cu(s)$?

a. $H_2|H^+\|Cu^{2+}|Cu$

b. $H^+|H_2\|Cu|Cu^{2+}$

c. $Cu^{2+}|Cu\|H_2|H^+$

d. $Cu|Cu^{2+}\|H^+|H_2$

19. When connected to a hydrogen electrode, an electrode with a negative standard reduction potential will carry out

a. reduction.

b. oxidation.

c. both oxidation and reduction.

d. neither oxidation nor reduction.

Use the table of standard reduction potentials below to answer the following questions.

Half-reaction	E^0 (volts)
$Al^{3+} + 3e^- \longrightarrow Al$	−1.662
$Ga^{3+} + 3e^- \longrightarrow Ga$	−0.549
$Tl^{3+} + 3e^- \longrightarrow Tl$	+0.741

20. Suppose you have two voltaic cells whose half-cells are represented by the following pairs of reduction half-reactions. For each voltaic cell, identify which half-reaction will proceed in the forward direction as a reduction and which will proceed in the reverse direction as an oxidation.

Voltaic Cell #1

$Al^{3+}(aq) + 3e^- \longrightarrow Al(s)$ ______________________

$Ga^{3+}(aq) + 3e^- \longrightarrow Ga(s)$ ______________________

Voltaic Cell #2

$Tl^{3+}(aq) + 3e^- \longrightarrow Tl(s)$ ______________________

$Ga^{3+}(aq) + 3e^- \longrightarrow Ga(s)$ ______________________

21. Calculate the cell standard potential, E^0_{cell}, of each voltaic cell in question 20.

Voltaic Cell #1: ____

Voltaic Cell #2:

Use the table of standard reduction potentials at the top of this page to answer the following questions.

22. Write the reduction and oxidation half-reactions for the following reaction:

$$Tl(s) + Al^{3+}(aq) \longrightarrow Tl^{3+}(aq) + Al(s)$$

reduction half-reaction: ____

oxidation half-reaction: ____

23. What is the standard reduction potential, E^0, for each half-reaction in question 22?

$E^0_{reduction}$: ______________________ $E^0_{oxidation}$: ______________________

24. Calculate the cell standard potential, E^0_{cell}, for the reaction in question 22.

____

25. Will the reaction in question 22 occur spontaneously as written? Explain why or why not.

____

26. Will the reverse reaction, $Tl^{3+}(aq) + Al(s) \longrightarrow Tl(s) + Al^{3+}(aq)$, occur spontaneously? Explain why or why not.

____

40. Write the balanced chemical equation for the standard cell notations listed below.

a. $I^- | I_2 \| Fe^{3+} | Fe^{2+}$

b. $Sn | Sn^{2+} \| Ag^+ | Ag$

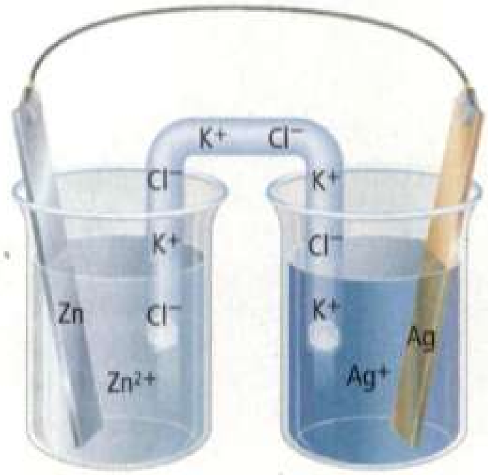
c. $Zn | Zn^{2+} \| Cd^{2+} | Cd$

41. Calculate the cell potentials for the following reactions.

a. $2Ag^+(aq) + Pb(s) \rightarrow Pb^{2+}(aq) + 2Ag(s)$

b. $Mn(s) + Ni^{2+}(aq) \rightarrow Mn^{2+}(aq) + Ni$

c. $I_2(aq) + Sn(s) \rightarrow 2I^-(aq) + Sn^{2+}(aq)$



■ Figure 25

42. **Figure 25** illustrates a voltaic cell consisting of a strip of zinc in a 1.0*M* solution of zinc nitrate and a strip of silver in a 1.0*M* solution of silver nitrate. Use the diagram and **Table 1** to answer these questions.

a. Identify the anode.
b. Identify the cathode.
c. Where does oxidation occur?
d. Where does reduction occur?
e. In which direction is the current flowing through the connecting wire?
f. In which direction are positive ions flowing through the salt bridge?
g. What is the cell potential at 25°C and 1 atm?

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67. Determine whether each redox reaction is spontaneous or nonspontaneous.

a. $Mn^{2+}(aq) + 2Br^{-}(aq) \rightarrow Br_2(l) + Mn(s)$

b. $2Fe^{2+}(aq) + Sn^{2+}(aq) \rightarrow 2Fe^{3+}(aq) + Sn(s)$

c. $Ni^{2+}(aq) + Mg(s) \rightarrow Mg^{2+}(aq) + Ni(s)$

d. $Pb^{2+}(aq) + 2Cu^{+}(aq) \rightarrow Pb(s) + 2Cu^{2+}(aq)$

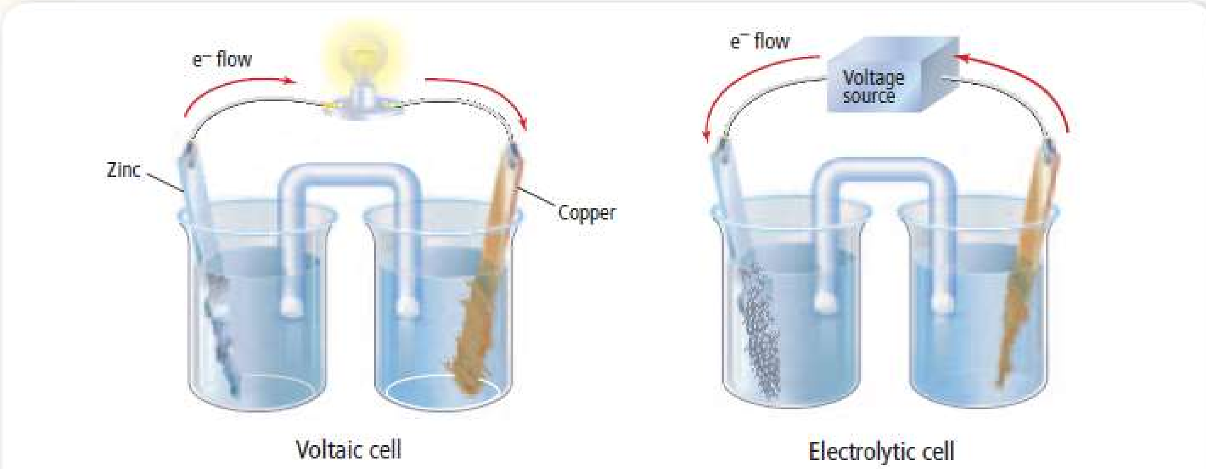
Section 3 : Electrolytes

Objectives - Compare between Voltaic cell and electrolytic cell

- Identify the products of water electrolysis

Reversing redox reactions:

- **Electrolysis:** the use of electrical energy to bring about a chemical reaction .
- **Electrolytic cell:** is an electrochemical cell in which electrolysis occurs , like the recharging of a secondary battery



In this voltaic cell, the oxidation of zinc supplies the electrons to light the bulb and reduce copper ions. The spontaneous reaction continues until the zinc is used up.

When an outside voltage is applied, the flow of electrons is reversed and the nonspontaneous reaction occurs, which restores the conditions of the cell.

	Voltaic cell	Electrolytic cell
The anode : The reaction:		
The cathode : The reaction:		
The charge of the anode :		
The charge of the cathode:		
The energy change:		
Is the reaction spontaneous?		

Objectives : - Identify the products of Molten NaCl electrolysis

Applications of Electrolysis:

1- Electrolysis of water: to generate hydrogen gas for commercial use.

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$$

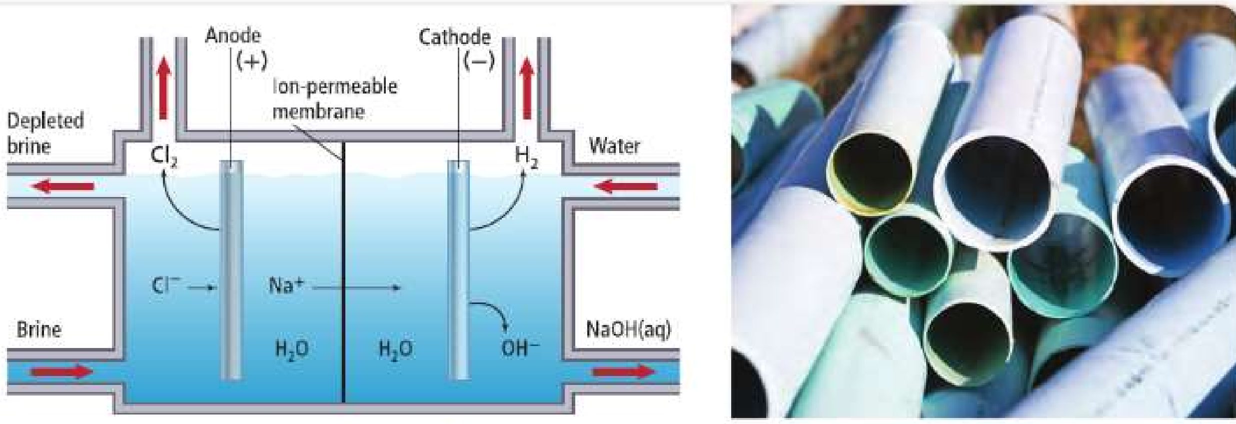
2- Electrolysis of molten NaCl :

- The name of the cell : **down's cell**
- The electrolyte : the molten NaCl
- The anode (+) made of carbon : ……………………………………………………………………
- The cathode (-) made of iron : ……………………………………………………………………
- The net cell reaction is : ………………………………………………………………………………

- The importance of chlorine :
  - to purify water for drinking and swimming
  - in cleaning products
  - in papers , plastics, insectisides, textiles , dyes and paints.
- The importance of sodium :
  - As a coolant in nuclear reactors
  - in sodium vapor lamps for outdoor lighting
  - to form ionic compounds (sodium salts) to use it in our foods .

Electrolytic cell	The product at the Anode (+)	The product at the Cathode (-)
1- Electrolysis of water		
2- **Electrolysis of molten NaCl (down's cell)**		
3- Electrolysis of brine (NaCl + H_2O)		

3- Electrolysis of brine:



Commercial facilities use an electrolytic process to obtain hydrogen gas, chlorine gas, and sodium hydroxide from brine.

Chlorine gas is used to manufacture polyvinyl chloride products, such as these pipes for water distribution.

Figure 21 In the electrolysis of brine (aqueous NaCl), sodium is not a product because water is easier to reduce.

- At the cathode:Tow reactions are possible : the reduction of sodium ions and the reduction of hydrogen in water molecules :

$$Na^+(aq) + e^- \rightarrow Na(s)$$

$$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$$

  - The reduction of water is easier than the reduction of sodium ions.
- At the anode: tow reactions are also possible: the oxidation of chlorine ions and the oxidation of oxygen in water molecules:

$$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$$

$$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$$

  - because the desired product is chlorine , the concetration of chlorine ions is kept high in order to favour this half – reaction.
- The overall cell reaction is :

$$2H_2O(l) + 2NaCl(aq) \rightarrow H_2(g) + Cl_2(g) + 2NaOH(aq)$$

- The three products are commercially importance subctances.

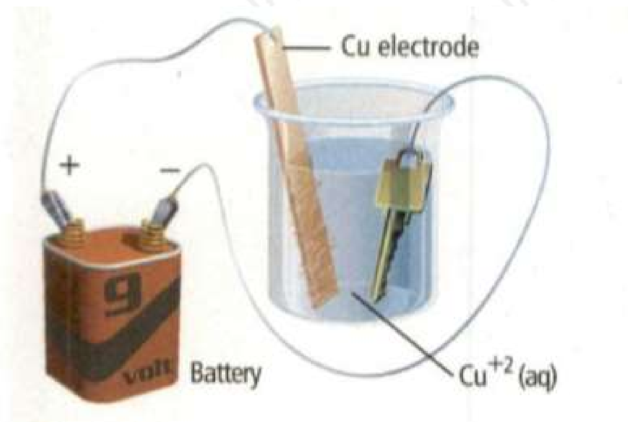
4- Electroplating :

- **Objects are** electroplated when a uniform coating is deposited usually as a protective or decorative layer.
- **The** cathode : is the object to be silver- plated for example, and silver ions are reduced to silver : ……………………………………………………………………………………
- **The** anode is the silver bar or sheet , silver is oxeidized to silver ions:

  ………
- **The** electrolyte is a solution of silver ions .
- **Current** passing through the cell must be carefully controlled in order to get a smooth , even metal coating.

Figure 23 Power is needed to oxidize silver at the anode and reduce silver at the cathode. In an electrolytic cell used for silver plating, the object to be plated is the cathode where silver ions in the electrolyte solution are reduced to silver metal and deposited on the object.

61. Electroplating **Figure 27** shows a key being electroplated with copper in an electrolytic cell. Where does oxidation occur? Explain your answer.

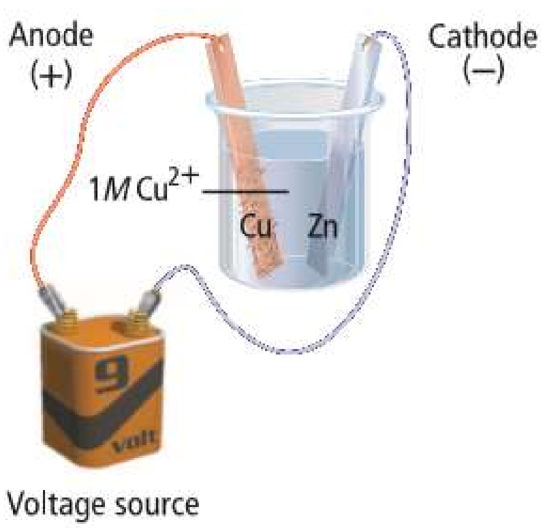


■ **Figure 27**

62. Answer the following questions based on **Figure 28.**

a. Which electrode grows? Write the reaction that occurs at this electrode.

b. Which electrode disappears? Write the reaction that occurs at this electrode.



■ **Figure 28**

63. Using **Figure 28,** explain what happens to the copper ions in solution.

....
....
....

65. Aluminum Production What substance is electrolyzed in the industrial process to produce aluminum metal?

....
....

72. Copper Refining In the electrolytic refining of copper, what factor determines which piece of copper is the anode and which is the cathode?

....
....

SECTION 3 REVIEW

22. Electrolysis is the process of using electrical energy to produce a chemical reaction. The electrolytic process is not spontaneous.

23. Electrolysis of brine involves an aqueous solution, which affects the products.

24. Cu atoms are oxidized to Cu 2+ then subsequently reduced to pure Cu atoms, with the impurities falling away.

25. The Hall-Héroult process requires high temperatures and a large amount of electricity to separate aluminum from its ore. Recycling requires only the heat needed to melt the metal.

26. The anode is a bar of gold; the cathode is the object to be plated.

27. First, a kilogram of silver contains many fewer atoms than a kilogram of aluminum because silver has a larger molar mass. Second, silver is easier to reduce. The reduction potential for silver is +0.7996 V. The reduction potential for aluminum is −1.662 V.

28. The Down's cell is a nonspontaneous reaction, so the potential should be negative. $E^0_{cell} = -4.07$ V

29. Student paragraphs should summarize the important ideas in the section. Refer to the Solutions Manual.

H.W : Solve questions 55 – 63 page 265 .

In the space at the left, write the word or phrase in parentheses that correctly completes the statement.

____________________ **1.** When a battery is being recharged, its redox reaction is reversed and energy is (absorbed, released) by the battery.

____________________ **2.** The use of electrical energy to cause a chemical reaction is called (combustion, electrolysis).

____________________ **3.** An electrochemical cell in which electrolysis is occurring is called an (electrolytic, exothermic) cell.

____________________ **4.** In a Down's cell, sodium metal and chlorine gas are produced from (molten, solid) sodium chloride.

____________________ **5.** The electrolysis of brine involves applying current to an aqueous solution of (hydrochloric acid, sodium chloride).

____________________ **6.** The commercially important products of the electrolysis of brine are hydrogen gas, chlorine gas, and (oxygen gas, sodium hydroxide).