MODULAR SYSTEM

METALS

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PREFACE

Chemistry is an interesting and fundamental science because it gives us a chance to explain the secrets of nature. Why does sodium burn in water? How do fireworks get flashing colors? Why carbon monoxide is poisonous but carbon dioxide is not?.... Many of these kind of questions and their answers are all part of world of elements. This book helps everyone to understand nature. However, one does not need to be a chemist or scientist to understand the simplicity within complexity around us.

There is no industry that does not depend on chemical substances: petroleum, pharmaceuticals, garment, aircraft, steel, and electronics, agricultural, etc. All need elements of different purity for their manufacture of thousands of substances.

Chemistry has its own language. To learn this language the first step is, of course, learning the letters in the chemistry alphabet, or the ELEMENTS. In this book either a single element or group of elements is studied. The aim was to write a modern, up-to-date book where students and teachers can get concise information about elements. Sometimes reactions are given in detailed form, but, in all, excessive detail has been omitted.

The book is designed to introduce basic knowledge about elements. Chemists work everyday to produce new compounds to make our lives easier with the help of this basic knowledge. In the design, emphasis has been placed on making this book student friendly. Throughout the books, colorful tables, important reactions, funny cartoons, interesting extras and reading passages are used to help explain ideas .This book will also show you how the elements are useful to us in everyday life. We hope, after studying world of elements, you will find chemistry in every part of your life.

The authors would like to thank Orhan Keskin and Ali Cavdar for their support and encouragement throughout the development of this book.

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We are particularly grateful to our wives and children for their patient during the writing of the book.

The Authors

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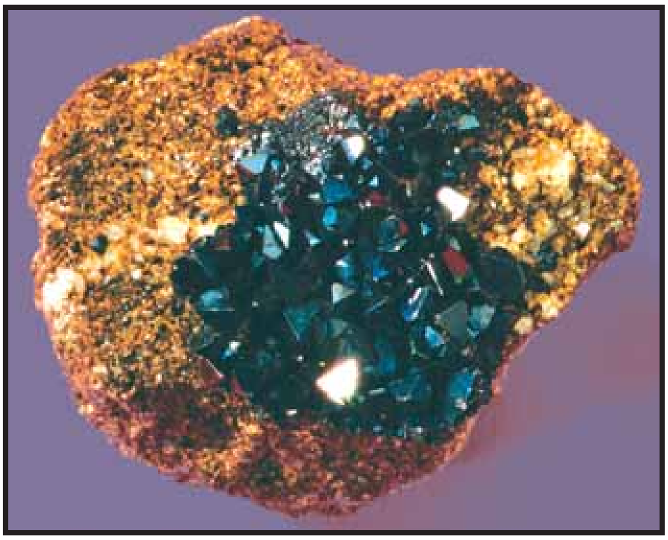
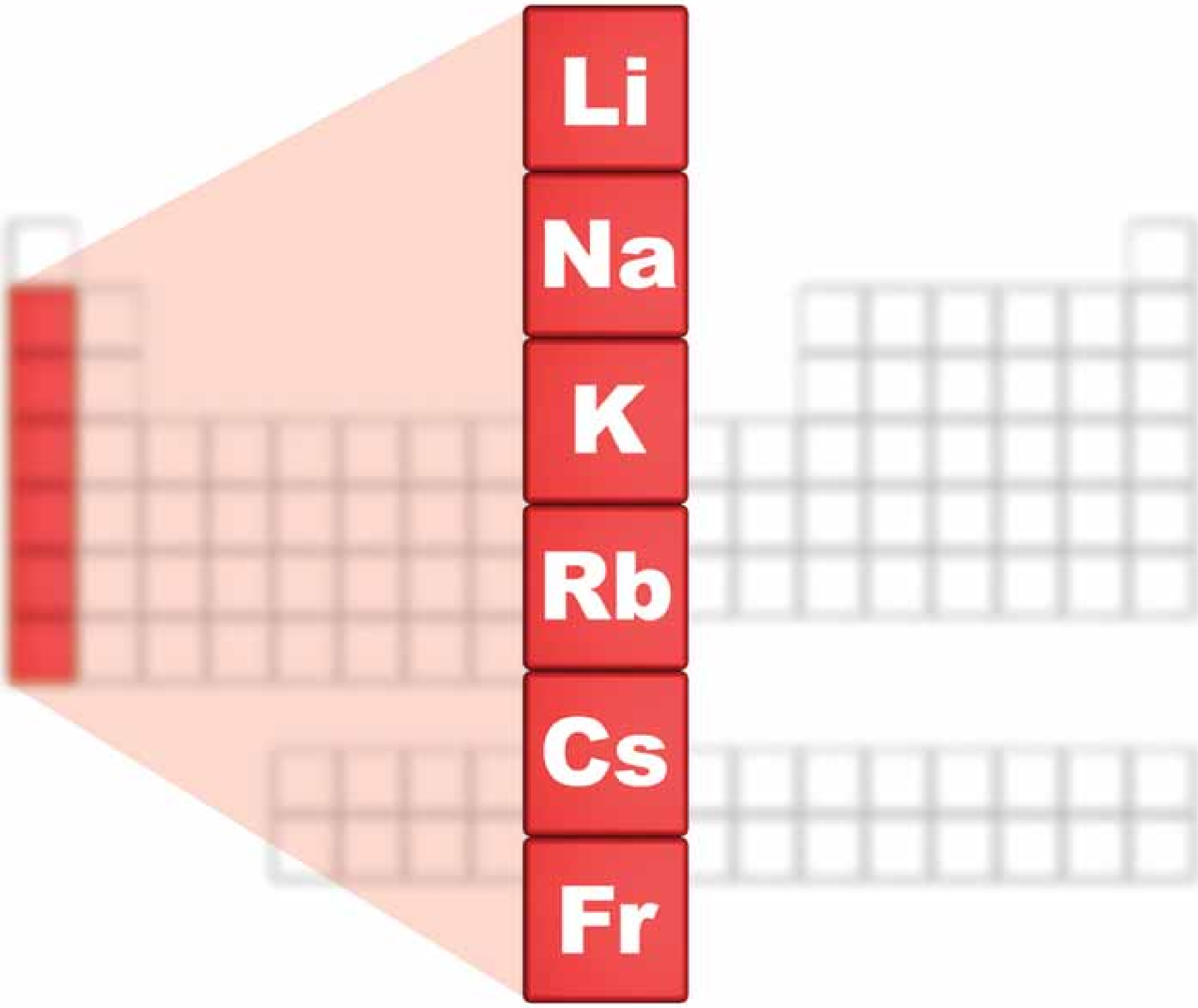
MODULAR SYSTEM

METALS



METALS



ALKALI METALS

INTRODUCTION

The elements of group 1A, except hydrogen, are called alkali metals. Although hydrogen is in group 1A, it shows nonmetallic properties.

The other members of the group, lithium (Li), sodium (Na), potassium (K), rubidium (Rb), cesium (Cs) and francium (Fr), show common metallic properties. Their electron configurations end with ns^1 and their atomic numbers are bigger than those of noble gases by 1. In other words, each period starts with an alkali metal, or each noble gas is followed by an alkali metal.

By giving their valence electron easily in chemical reactions, they form +1 charged ions. Because the alkali metals are the elements which have the least ionization energy and the highest atomic radius, in each period they are a group of most active metals. Since the activity of metal increases from top to bottom, francium, Fr, is expected to be the most active metal in the periodic table. In reality, the element cesium, Cs, is the most active metal because francium is a radioactive element (has been isolated only in minute quantities) and its properties have not been determined at all. Some properties of alkali metals are given in the table below.

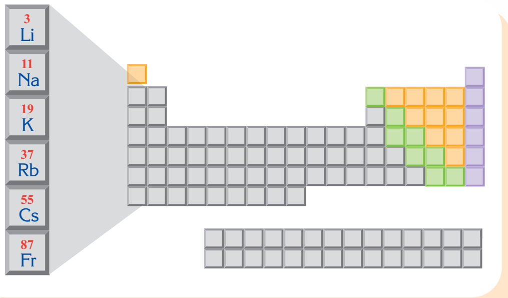


Figure 1 The alkali metals (group 1A)

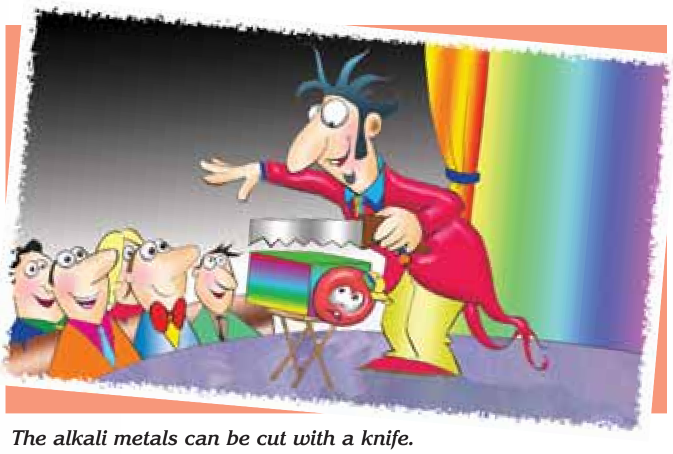
The name "alkali" comes from the Arabic word "El-kâli" meaning basic.

Name	Lithium	Sodium	Potassium
Symbol	Li	Na	K
Atomic number	3	11	19
Atomic mass	6.94	22.99	39.10
Electron configuration	$[He]2s^1$	$[Ne]3s^1$	$[Ar]4s^1$
Melting point (°C)	180.5	97.7	63.4
Boiling point (°C)	1342	883	759
Density (g/cm³)	0.54	0.97	0.86
Ist Ionization energy (kj/mol)	520.2	495.8	418.8
Atomic radius (pm)	134	154	196
Common oxidation numbers	+1	+1	+1
Color	silvery white	silvery white	silvery white
Physical state at 25°C	solid	solid	solid
Origin and meaning of name	*lithos* - stone	*natrium* - soda	*potash* - kalium
Earth's crust abundance (%)	1.7×10^{-3}	2.3	1.5

Table 1 Some properties of alkali metals.

Since the electron configurations of alkali metal ions with +1 charge are the same as those of noble gases, forming +2 charged ions by giving the second electron is very difficult for them. Therefore, in order to take the second electron of an alkali metal, very high energy is needed. The highest second ionization energy belongs to the alkali metals in the same period.

The melting points, boiling points and densities of alkali metals are lower than those of other metals. The densities of lithium, sodium and potassium are lower than water. As the atomic number of alkali metals increases, their melting and boiling points and their densities decrease. They are rather soft (even lithium, the hardest metal in the group) and can easily be cut with a kitchen knife. Their newly cut surfaces are silvery grey. They are good conductors of heat and electricity.



The alkali metals can be cut with a knife.

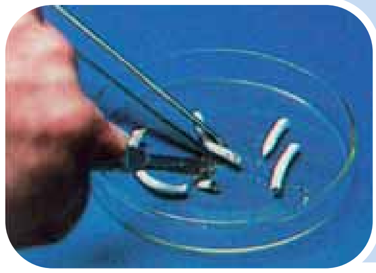


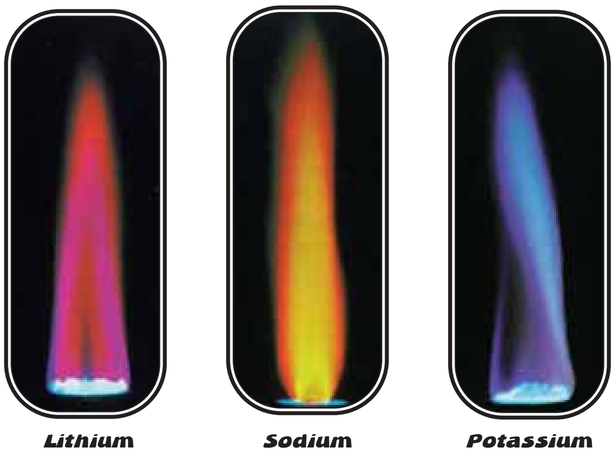
Figure 2 *Even lithium, the hardest alkali metal, may be cut with a knife.*

	Name	Formula
Li	ores petalite	$Li_3Al(PO_4)_2$
	spodumene	$LiAl(Si_2O_3)_2$
	amblygonite	$LiAl(F, OH)PO_4$
Na	chile saltpeter	$NaNO_3$
	washing soda	Na_2CO_3
	baking soda	$NaHCO_3$
	kryolite	Na_3AlF_6
	borax	$Na_2B_4O_7 . 10H_2O$
	sodium sulfate	Na_2SO_4
	albite	$NaAlSi_3O_8$
K	sylvite	KCl
	carnallite	$KCl \cdot MgCl_2 \cdot 6H_2O$
	orthoclas	$KAlSi_3O_8$
	mica	$KH_2Al_3(SiO_4)_3$
	potassium	K_2SO_4
	fertilisers	KNO_3

Table 2 *Mineral of Li, Na and K.*

Even at normal conditions they are found in monoatomic structure. At high temperatures their diatomic molecules can be observed.

When the aqueous solutions of their +1 charged ions are exposed to flame, they give a characteristic flame color: lithium-red, sodium-yellow, potassium and rubidium-violet, and cesium-blue.



Flame tests of alkali metals

1. OCCURRENCE

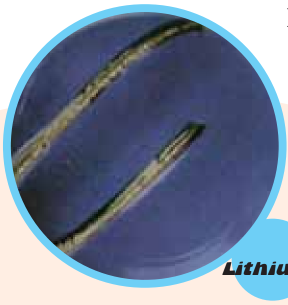
Since the alkali metals are the most active metals, they are not found free in nature, but as compounds. Generally they occur in sea water, in some lakes, in rock salt sources and in the compounds of oxygen and silicon as silicates.

Sodium and potassium are present in vast amounts in nature. Lithium is less abundant in comparison with sodium and potassium. Rubidium and cesium are found in trace amount in nature. Francium, a radioactive element, is a side product and all isotopes of francium are also radioactive. The most stable isotope of francium is ^{223}Fr, with a half-life of 23 minutes.

Now let's focus on lithium, sodium and potassium.

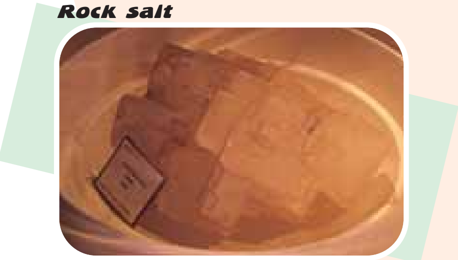
Lithium, Li

Lithium is found in nature in the ores petalite ($Li(AlSi_4)O_{10}$), spodumene ($LiAl(Si_2O_3)_2$) amblygonite and lepidolite.


Lithium


Sodium


Potassium


Rock salt

Only specialists try to remember complicated formulas, such as those of petalite, spodumene, lepidolite and etc.

Sodium, Na

The most important compound of sodium is sodium chloride, NaCl. It is found in the form of NaCl in the seas and in rock-salt layers formed by the evaporation of ancient lakes and seas. The important sodium sources are Chile saltpeter ($NaNO_3$), washing soda (Na_2CO_3) and baking soda ($NaHCO_3$). The most important sodium minerals are kryolite (Na_3AlF_6), borax ($Na_2B_4O_7 . 10H_2O$), sodium sulfate (Na_2SO_4) and albite ($NaAlSi_3O_8$).

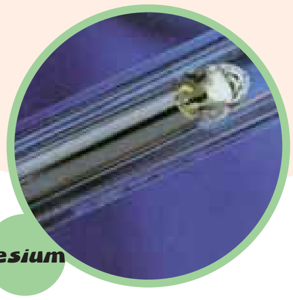

Borax

Potassium, K

Potassium constitutes 1.5% of the earth's crust. Potassium is found as the minerals sylvite (KCl) and carnallite ($KCl \cdot MgCl_2 \cdot 6H_2O$) and as the silicates of orthoclas ($KAlSi_3O_8$) and mica ($KH_2Al_3(SiO_4)_3$). The main sources of potassium are K_2SO_4 and KNO_3, which are used as fertilizers.


Rubidium


Mica


Cesium


Francium

Preparation of Sodium

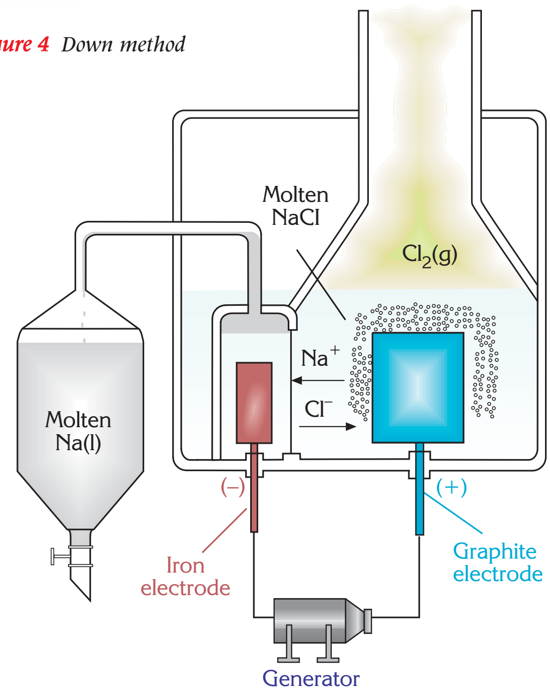
The alkali metals are made by electrolysis of molten hydroxides or halides. Because of their reactivity, the metals must be kept in an inert atmosphere or under oil. The preparation of sodium is the best example of the preparation of alkali metals.

Figure 3 *Sir Humphry Davy (1778-1829)*

Elemental sodium was first obtained in 1807, by Davy, by the electrolysis of molten NaOH. Today molten NaCl is used instead of molten NaOH. This method is called the Down method shown in Figure 4.

The melting point of NaCl is 800°C. However, if a quantity of $CaCl_2$ three times as much as NaCl is added, this temperature is easily reduced to about 600°C. In the Down device, graphite is used as an anode, and iron, steel or copper is used as a cathode.

Figure 4 *Down method*



As the Cl^- ions are evolved at the anode as chlorine gas, the Na^+ ions are deposite at the cathode as metallic sodium, Na.

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

Cathode $2Na^+ + 2e^- \rightarrow 2Na(l)$

With the help of a funnel surrounding the anode, the chlorine gas is collected outside the system. A small quantity of Ca is deposited at the cathode. Since calcium rarely reacts with molten sodium, and because it is denser than sodium, it is easily separated from the medium. The liquid sodium obtained is less dense than the electrolyte, so it accumulates on the surface, and through a small pipe it can be collected in the funnel. The sodium and chlorine produced can react very vigorously with each other. In order to avoid such a reaction, the cathode is surrounded by a diaphragm made from an iron bar.

2. CHEMICAL PROPERTIES

Alkali metals are very strong reducing agents. They all react with cold water. Cs and Rb violently react with water. Sodium, potassium and lithium become dull when they contact air. Therefore, they are kept in kerosene or paraffins to prevent reactions with oxygen.

Since they are alkali, they do not react with bases but react with acids. Their compounds with group 7A elements are called salts. Whereas the halo-compounds of hydrogen have acidic properties, the oxides of alkali metals are basic, and the aqueous solution of their oxides give bases.

The exothermic reaction of potassium with water.

Reactions

1. Alkali metals are strong reducing agents.

$nM^0 + Y^{+n} \longrightarrow nM^+ + Y^0(s)$

(Y : any metal with low activity,
M : any alkali metal)

$6\overset{0}{Na}(s) + \overset{+3}{Al_2}O_3(s) \longrightarrow 3\overset{+1}{Na_2}O(s) + 2\overset{0}{Al}(s)$

$3\overset{0}{K}(s) + \overset{+3}{Cr}Cl_3(s) \longrightarrow 3\overset{+1}{K}Cl(s) + \overset{0}{Cr}(s)$

2. Hydrides are formed as a result of alkali metals' reactions with hydrogen. Hydrides contain a +1 charged alkali metal and –1 charged hydrogen.

$2M^0(s) + H_2{}^0(g) \xrightarrow{heat} 2M^+H^-(s)$

(M : any alkali metal)

$2K^0(s) + H_2{}^0(g) \longrightarrow \underset{\text{potassium hydride}}{2KH(s)}$

$2Na(s) + H_2(g) \longrightarrow \underset{\text{sodium hydride}}{2NaH(s)}$

The hydrides obtained react with water to form solutions of alkali hydroxides and H_2 gas.

$MH(s) + H_2O(l) \longrightarrow M^+OH^-(aq) + H_2(g)$

(MH : an alkali hydride, MOH : an alkali hydroxide)

$KH(s) + H_2O(l) \longrightarrow KOH(aq) + H_2(g)$

$NaH(s) + H_2O(l) \longrightarrow NaOH(aq) + H_2(g)$

!

ATTENTION

Flammable hydrogen gas is produced by the vigorous reaction of K and Na with H_2O.

3. They react with water violently. As a result of this reaction H_2 gas and a base solution form.

$2M(s) + 2H_2O(l) \rightarrow 2MOH(aq) + H_2(g) + heat$

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g) + heat$ (Figure 5).

The hydroxides formed are thermally stable, except LiOH.

$2LiOH \longrightarrow Li_2O + H_2O$

4. They may form oxides, peroxides or superoxides by reacting with oxygen in the air.

As a result of reactions with excess oxygen, lithium forms oxide, Li_2O, sodium forms peroxide Na_2O_2 and potassium, rubidium and cesium form superoxides, such as KO_2, RbO_2, CsO_2.

$4Li(s) + O_2(g) \longrightarrow 2Li_2O(s)$

$2Na(s) + O_2(g) \longrightarrow Na_2O_2(s)$

$K(s) + O_2(g) \longrightarrow KO_2(s)$

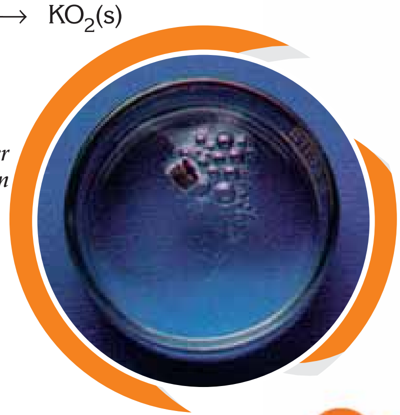


Figure 5 *Na floats on water and produces H_2 gas when added to water.*

To obtain the oxides of sodium and potassium (Na_2O and K_2O), their peroxides and superoxides are affected by their own metals.

$Na_2O_2 + 2Na \rightarrow 2Na_2O$ $\quad$ $KO_2 + 3K \rightarrow 2K_2O$

5. All of them react with halogens to form alkali halides (salts of alkali metals).

   $2M(s) + X_2(g) \longrightarrow 2MX(s)$ (Figure 6).

   (M : alkali metal, X : halogen,

   MX ; alkali metal halides)

   $2Li(s) + F_2(g) \longrightarrow 2LiF(s)$

   $2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$

Even though H takes +1 charge, it takes –1 with metals.

6. They do not react with bases

   $M(s) + OH^-(aq) \longrightarrow$ No reaction

7. When they react with acids, they produce salts and liberate H_2 gas.

   $M(s) + HX(aq) \longrightarrow MX(aq) + 1/2H_2(g)$

   (HX : Halo acid, MX : salt of an alkali)

   $2K(s) + 2HCl(aq) \longrightarrow 2KCl(aq) + H_2(g)$

   $2Na(s) + 2HBr(aq) \longrightarrow 2NaBr(aq) + H_2(g)$

Figure 6 *Potassium reacts with chlorine gas vigorously to form potassium chloride salt.*

$2K(s) + Cl_2(g) \longrightarrow 2KCl(s)$

3. COMPOUNDS

Generally the compounds are typically ionic. They have high melting and boiling points and dissolve better in water than in non-polar solvents.

Appearance of salt crystals under a microscope.

The compounds of lithium are the exception, and show a good deal of covalency. The reason for this can be explained by lithium's very small atomic size and its outer electron's proximity to the nucleus. For example, the alkali metal salts are very soluble in water. However, some salts of Li (Li_3PO_4, Li_2CO_3 and LiF) are only slightly soluble in water.

1. Halides

Now the most important halide of alkali metals, sodium chloride, will be considered.

Sodium chloride, NaCl

NaCl, known as table salt, melts at 801°C and boils at 1465°C. It is a combination of colorless, transparent, cubic structured crystals.

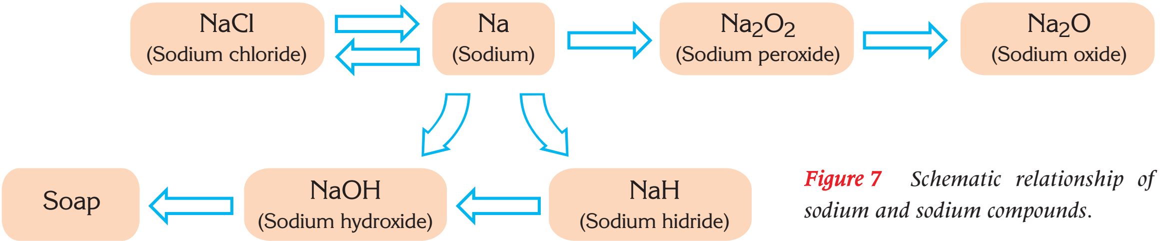


Figure 7 *Schematic relationship of sodium and sodium compounds.*

NaCl is found as rock salt in nature, and dissolved in sea water at a ratio of 3%. NaCl has a characteristic salt taste and can be separated from sea water with 96% purity. For this process, sea water is put into large-surfaced pools. Then water is slowly evaporated. As a result, NaCl is obtained in crystalline structures.

NaCl is an ionic compound. There are electrostatic attraction forces between its ions, Na^+ and Cl^-. NaCl dissolves in water as follow.

$NaCl(aq) \rightarrow Na^+(aq) + Cl^-(aq)$

This dissolution process occurs when H_2O (water) dipoles surround the ions found at the surface of the ionic crystals. This is called dehydration (Figure 8).

$Na^+ + 6H_2O \rightarrow Na(H_2O)_6^{\ +}$

Its solubility in water changes slightly with temperature. For example, 35.7 grams of salt dissolves in $100cm^3$ water at 0°C, while 39.1 grams of salt dissolves in $100cm^3$ water at 100°C.

A solution of NaCl can be used in the production of NaOH with the electrolysis method. By-products of the process are hydrogen and chlorine gases.

$$2NaCl + 2H_2O \rightarrow \underset{\text{cathode}}{H_2(g)} + \underset{\text{anode}}{Cl_2(g)} + 2NaOH(aq)$$

When molten NaCl is electrolyzed, liquid Na and Cl_2 gas are obtained.

$$2NaCl(l) \rightarrow \underset{\text{cathode}}{2Na(l)} + \underset{\text{anode}}{Cl_2(g)}$$

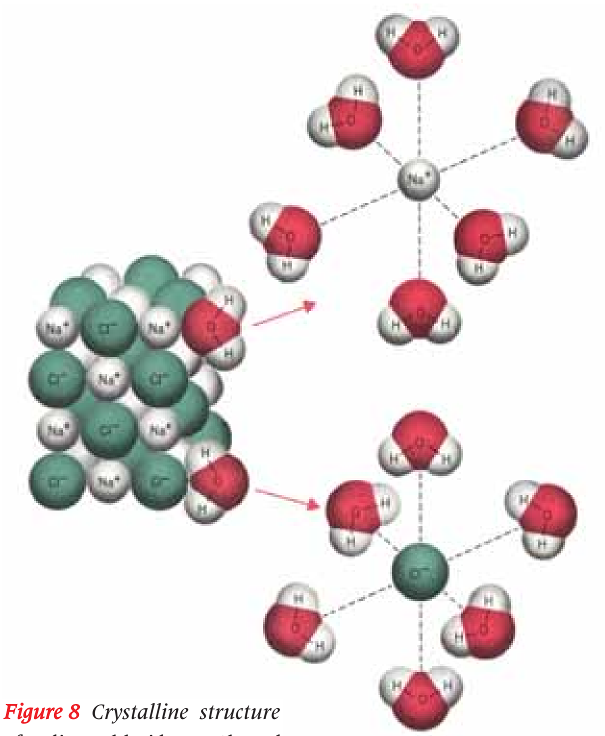


Figure 8 *Crystalline structure of sodium chloride, NaCl, and its dehydration.*

A NaCl solution containing NaCl(s) in a ratio of 23.6% by mass freezes at –23°C. During winter, rock salt is spread on streets, pavement and highways in order to melt the ice.

Table salt also has a property that prevents food from spoiling. Before refrigerators, salting was an extremely useful method of food preservation. Cooling is used today.

Salt is added to food to give flavor instead of preserving it. For example, to butter, which can be protected for a long time in refrigerators, 2% table salt is added for flavor. Even though it is not necessary to keep margarine in refrigerators, table salt is also added for the same reason. Some foods like olives, cheese and pickles contain abundant salt.

Blood plasma and other body fluids contain about 0.9 grams of sodium chloride per 100 mL.The amount of salt needed by an adult is 0.5 grams per day. But generally, people use much more salt than they need. As a result, illnesses like hypertension and kidney deficiencies arise.

NaCl is the main substance in the preparation of other salts of sodium. For example, NaCl, a neutral salt, reacts with concentrated H_2SO_4 and produces Na_2SO_4 or $NaHSO_4$ salt.

$$2NaCl + H_2SO_4 \longrightarrow Na_2SO_4 + 2HCl$$

$$NaCl + H_2SO_4 \longrightarrow NaHSO_4 + HCl$$

Recent studies have shown that there is a striking relation ship between salt use and hypertension. Hypertension can be reduced without using medicine simply by decreasing salt use.

Children should be taught to use only a small amount of salt. It is significant that in many countries baby foods are prepared free of salt. We have to ensure our adolescents not to consume fast foods with a high salt content, so that we can decrease the risk of illness.

2. Hydroxides

The hydroxides of alkali metals are all white, hydroscopic, crystalline ionic solids containing the metal ion, M^+, and the hydroxide ion, OH^-.

They dissolve in water giving metal and hydroxide ions (M^+ and OH^-).

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

Because of the high ionization percent in water, they behave as strong bases and react with acidic and amphoteric substances.

$$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$$

$$NaOH(aq) + NH_4Cl(aq) \rightarrow NaCl(aq) + NH_3(g) + H_2O(l)$$

$$2NaOH(aq) + CO_2(g) \rightarrow Na_2CO_3(aq) + H_2O(l)$$

$$6KOH(aq) + Al_2O_3(s) \rightarrow 2K_3AlO_3(aq) + 3H_2O(l)$$

NaOH and KOH are used to remove CO_2 and SO_2 from air.

If exposed to air, they absorb water vapor very readily and form concentrated aqueous solutions of the hydroxides, which are all deliquescent.

Deliquescence is defined as the absorption of moisture from the atmosphere by a solid to form a solution.

Sodium hydroxide and potassium hydroxide have important uses in the manufacture of soap and detergents (Figure 9). In industry and commerce they are called caustic soda and caustic potash respectively. They have other applications, such as in the production of many chemicals, textiles, cleanser, paper, pulp and in petroleum refining.

Figure 9 *NaOH is used in the preparation of solid soap whereas KOH is used in the preparation of soap like gel.*

3. Nitrates

Sodium nitrate, $NaNO_3$ (Chile saltpeter), is found in large deposits in Chile. It decomposes by releasing oxygen gas at about 500°C.

$$2NaNO_3(s) \xrightarrow{\text{heat}} 2NaNO_2(s) + O_2(g)$$

Potassium nitrate, KNO_3 (saltpeter) is also decomposed by heating.

$$2KNO_3(s) \xrightarrow{\text{heat}} 2KNO_2(s) + O_2(g)$$

KNO_3 is prepared by the reaction of KCl and $NaNO_3$ under 100°C.

$$KCl(aq) + NaNO_3(aq) \longrightarrow NaCl(aq) + KNO_3(aq)$$

Lithium and its compounds show some differences from other alkali metals and compounds.

Formula	Common Name	Uses
LiOH		Synthesis of organic compounds
Li_2CO_3		Ceramics, glasses
NaOH	Caustic soda	Soap, paper, textile, dye
NaCl	Table salt	Nutrition, soap, glass, pottery
Na_2SO_4	Glauber's salt	Glass, dye, medicine
$NaHCO_3$	Baking soda	Baking powder, neutralization of HCl in stomach
Na_2CO_3	Washing soda	Cleaning
$NaNO_3$	Chile saltpeter	Fertilizer
KOH	Potash	Soap
KCl	Sylvite	Fertilizer
K_2SO_4		Fertilizer
K_2CO_3	Pearl ash	Production of glass and soap
KBr		Medicine, glue, photography
KI		Tincture
KNO_3	Saltpeter	Gunpowder, fertilizer

Table 3 *Common names and uses of some alkali metal compounds.*

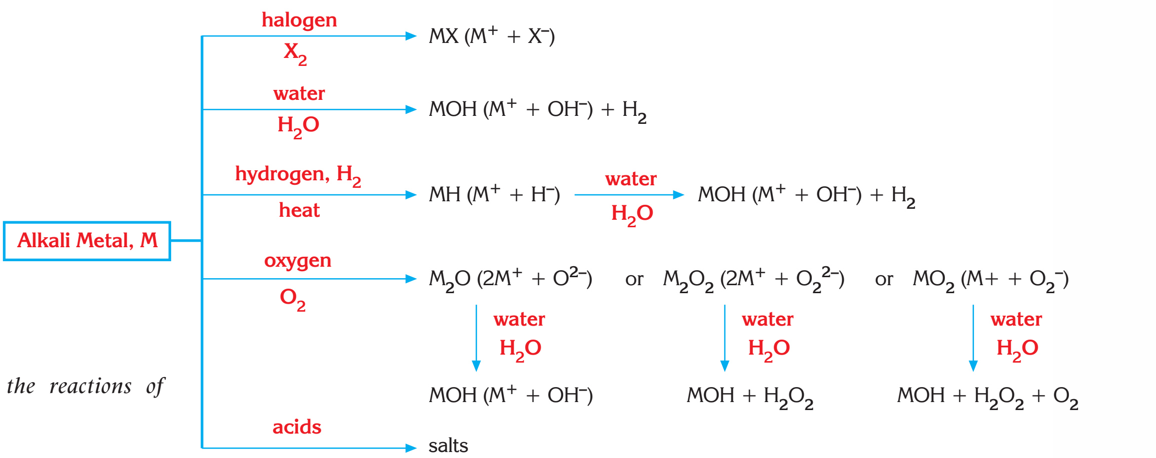


Figure 10

Summary of the reactions of alkali metals.

KNO_3 is the main component of gunpowder, with sulfur and wood charcoal. When gunpowder is heated, the following reaction occurs.

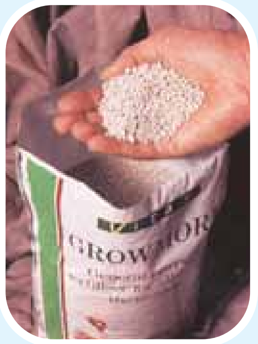
$$2KNO_3(s) + S(s) + 3C(s) \rightarrow K_2S(s) + N_2(g) + 3CO_2(g)$$

The sudden formation of hot N_2 and CO_2 gases causes an explosion.

Lithium nitrate, $LiNO_3$, shows differences in thermal decomposition.

$$4LiNO_3 \longrightarrow 2Li_2O + 4NO_2 + O_2$$

USES

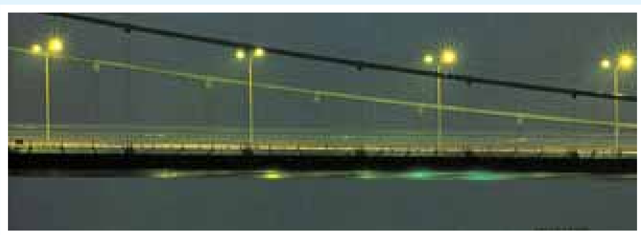


Lithium is used to eliminate poisonous gases from copper and copper alloys. Some lithium compounds are used as medicine, such as Li_2CO_3 in psychiatry and $LiHCO_3$ in the treatment of hyperureacemia. LiCl is a drying agent and $LiAlH_4$ is used as a reducing agent in organic reactions.

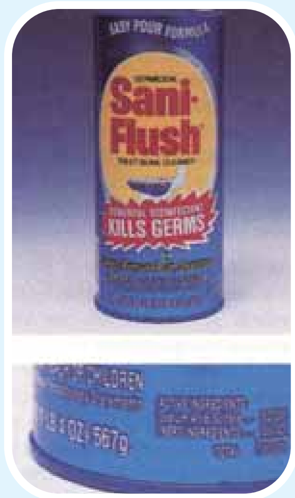
Sodium is used to produce electricity in nuclear reactors by transferring excess heat to the vapor turbines. In metallurgy, it is used to produce pure metals from their oxides. It is also used in sodium vapor lamps for foggy days. Its salts are used in the medical industry. In industry, it has an important role in the production of salts, soaps, baking soda, glass and pigment.

Potassium is mostly used in fertilizers, producing soaps, and medicine in several areas. The soaps in which potassium is used are very soft and have the common name gel.

Since rubidium and cesium emit electrons by photocell rays, they are used as photocells. They are used in TV receivers, as a prism in spectrophotometers, and as propellants in space ships. In these processes, cesium is used rather than rubidium.



Sodium lamps are used in street lamps and in cars. When it is foggy, they emit yellow light.



SUPPLEMENTARY QUESTIONS

1. Write the names of the alkali metals with their symbols.

2. Why are alkali metals located in group 1A in the periodic table?

3. Why is it difficult to take the second electron from alkali metals?

4. Alkali metals can not be found freely in nature.

a. What is the main reason?

b. Write down briefly where sodium metal is found in nature.

5. Write the common names of ores that are natural sources of alkali metals.

a. Na_2CO_3 ....

b. KCl ....

c. $NaNO_3$ ....

d. NaOH ....

6. Write the correct formulas of the following.

a) Table salt b) Borax c) Washing soda d) Gunpowder

7. Fill in the blanks with the appropriate alkali metal.

a. The characteristic color of .... is yellow, ... is red and .... blue in the flame test.

b. .... has the highest ionization energy.

c. .... has the biggest radius.

d. .... is the most reactive.

e. .... has the highest melting and boiling point.

f. .... is the least dense.

g. .... is radioactive.

8. Find the mass percentages of sodium in the following compounds.

a. $NaHCO_3$ **b.** Na_3AlF_6

c. $Na_2CO_3 . 10H_2O$ **d.** $Na_2B_4O_7 . 10H_2O$

9. ${}_{11}Na \rightarrow {}_{11}Na^+ + e^-$

${}_{10}Ne \rightarrow {}_{10}Ne^+ + e^-$

Which of the above reactions needs more energy? Explain.

10. Complete and balance the following equations.

a. $NaH(s) + H_2O(1) \longrightarrow$ .... + ...

b. $6Na(s) + Al_2O_3(s) \xrightarrow{heat}$ .... + ...

c. $Li(s) + H_2O(l) \longrightarrow$ .... + ...

d. $2Na(s) + O_2(g) \longrightarrow$ .... + ...

e. $Li(s) + F_2(g) \longrightarrow$ .... + ...

f. $K(s) + HCl(aq) \longrightarrow$ .... + ...

11. What are the uses of sodium metal?

12. What are the uses of lithium metal?

13. If we use 1.38 g of Li metal to remove CO_2 gas from a room with dimensions 8 m x 8 m x 5.75 m, what is the percentage of CO_2 gas in the room?

$4Li + 2CO_2 + 4H_2O \longrightarrow Li_2O + Li_2CO_3 + C$

14. The chemical reaction is

$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$

during the inflation of air bags. If a 6 cm^3 of air bag is blown up at 1 atm and 0°C, what is the mass of Na produced?

15. Write the reaction equations for the following reaction schemas.

a. $NaCl \rightarrow Na \rightarrow NaH \rightarrow NaOH \rightarrow Na_2SO_4$

b. $LiCl \rightarrow Li \rightarrow Li_2O \rightarrow LiCl$

c. $K \rightarrow KOH \rightarrow K_2SO_4$

16. Write the products of electrolysis of the given substances.

a. Molten potassium bromide

b. Lithium chloride solution

c. Molten potassium hydroxide

17. During the electrolysis of NaCl solution, a total of 13.44 L gas is obtained at STP. According to the given informations, find the masses of each substance produced in this process.

18. If 67.2 L SO_3 gas (at STP) is completely dissolved in 25% NaOH solution (1 kg), find the mass percentages of all species in the final solution.

19. When 10.55 g of the mixture of Glauberg salt ($Na_2SO_4 . 10H_2O$) and washing soda ($Na_2CO_3 . 10H_2O$) is heated, 4.25 g H_2O is obtained. Find the mass percentage of washing soda in the given mixture.

20. When the mixture of lithium hydride and sodium hydride is added to 193 mL pure water, the total mass of all substances decreases 1 g. The mass percentage of bases is detected as 8% after the reaction. Find the mole numbers of the hydrides used.

MULTIPLE CHOICE QUESTIONS

1. Which one of the following is not a member of the alkali metals group?

A) Li B) Ba C) Na D) K E) Rb

2. Which one of the following atoms has the smallest volume?

A) Li B) Na C) K D) Rb E) Cs

3. What is the color of sodium ion in the flame test?

A) Blue B) Green C) Red D) White E) Yellow

4. Which one of the following metals is radioactive?

A) Li B) Na C) Fr D) K E) Cs

5. What is the common charge of alkali metals?

A) +1 B) +2 C) +3 D) +4 E) –1

6. What is the number of electrons in the outermost shell of the atoms in the alkali metals group?

A) 8 B) 5 C) 4 D) 3 E) 1

7. What are the oxidation states of the nonmetals for the given compounds, respectively: NaH, K_2O, Li_3N?

A) –1, +2, +3 B) –1, +1, +2 C) +1, –2, –3 D) –1, –2, –3 E) –1, –2, +3

8. Which one of the following metals is the most active?

A) Rb B) Fe C) Cs D) Li E) Na

9. Which one of the following compounds of sodium is the most widely used in daily life?

A) Na_2SO_4 B) NaBr C) NaCl D) NaI E) $NaNO_3$

10. Why do Li, Na and K float in water?

A) Low areas B) Volumes C) Low densities D) Masses E) Physical states

11. How are alkali metals kept from reacting with O_2 in the air?

A) In water B) In base C) In acid D) In alcohol E) In kerosene

12. What is the charge of oxygen in the compound produced at the end of the reaction?

$Na + Al_2O_3 \longrightarrow$

A) –1 B) –2 C) +2 D) +1 E) 0

13. When 0.7 g of Lithium react with water how many grams of $H_2(g)$ are produced?

A) 0.1 B) 0.2 C) 0.3 D) 0.4 E) 0.5

14. When 4.6 g of sodium react with O_2, how many grams of Na_2O_2 are produced?

A) 18.6 B) 24.8 C) 12.4 D) 7.8 E) 6.2

15. If 7.45 g of KCl is produced at the end of the reaction of potassium with chlorine, how many grams of K are consumed?

A) 3.5 B) 1.95 C) 15.6 D) 3.9 E) 7.8

16. What is the mass of H_2 gas produced at the end of the reaction of 1.4 g of Li with 40.5 g of HBr?

A) 0.2 g B) 10.4 g C) 0.8 g D) 1.6 g E) 3.2 g

17. Which metal salt is used in the production of soap?

A) Li B) Rb C) Na D) Cs E) Fr

18. Who first found elemental sodium?

A) Lussac B) Davy C) Down D) Mendeleyev E) Boyle

19. Which one of the following is used in making tincture?

A) LiCl B) CaI C) NaI D) RbF E) KI

20. What is the name of the method used to produce sodium today?

A) Hess B) Davy C) Down D) Humphry E) Bohr

21. I. Electrolysis of molten NaCl

II. Electrolysis of NaCl solution

III. Electrolysis of NaOH solution

Which of the above is/are used to produce metallic sodium in industry?

A) I Only B) II Only C) I and II D) I and III E) I, II and III

PUZZLE

The words below are listed with their descriptions, but one letter is missing from each word. List the missing letters. When the letters are read downwards they will spell the surnames of the three scientists whose work was involved in the classification of elements. A clue for each scientist is given.

1. He looked at the patterns between the relative atomic masses of elements and their densities.

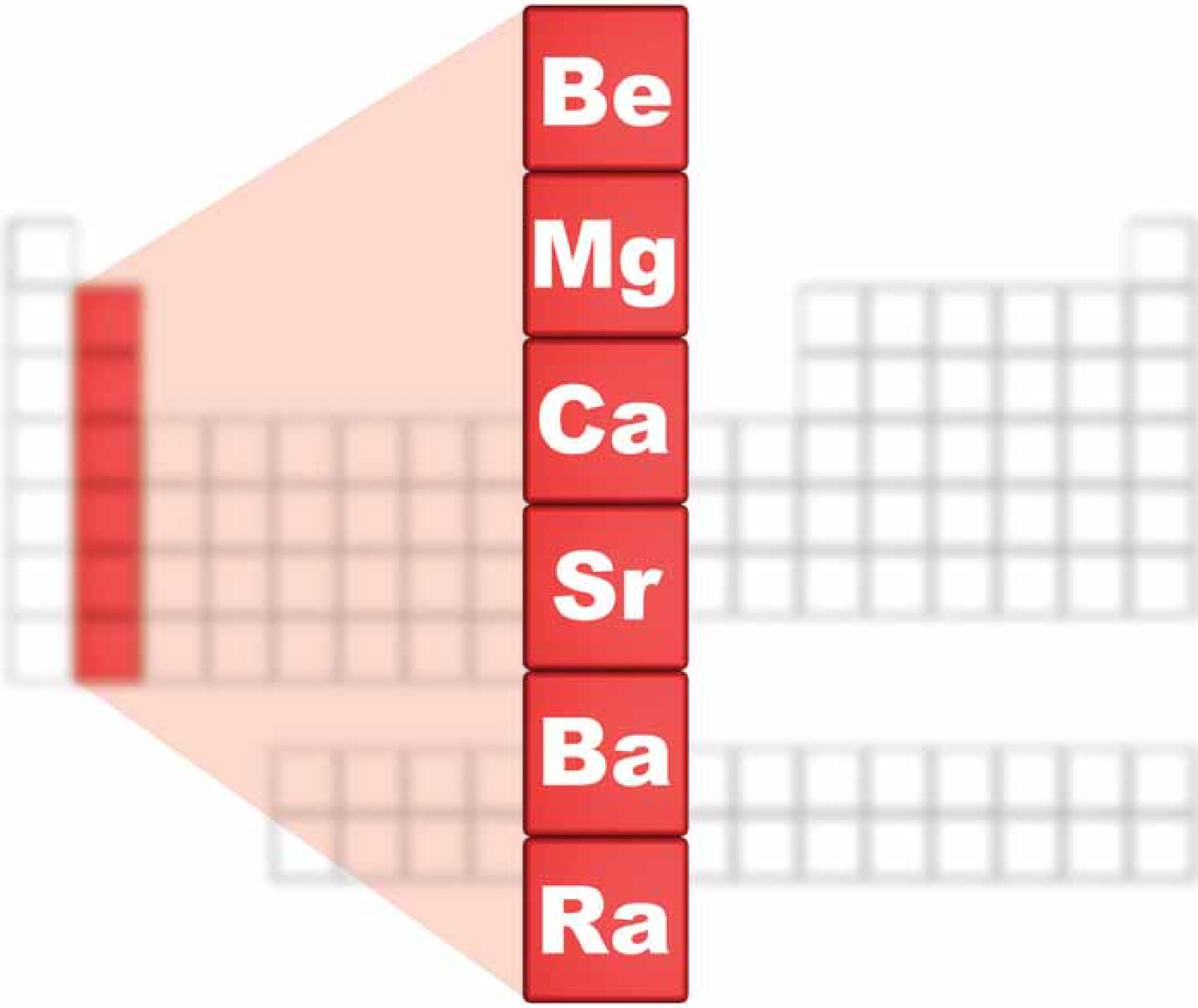
_______	DISPLACEENT	Cholorine's oxidation of bromine is an example of this	_______
_______	CSIUM	The most active metal after Fr.	_______
_______	SHIN	Alkali metals are this when freshly cut.	_______
_______	PRIODS	Horizontal rows of elements in the periodic table	_______
_______	AI	A mixture of gases including oxygen, nitrogen, carbon dioxide and inert gases	_______

2. When he drew up his periodic table he left gaps for elements he predicted must exist but had not yet been discovered.

_______	POTASSIU	Fourth most reactive alkali metal.	_______
_______	COLOURD	Transition elements form compounds which are this.	_______
_______	REACTIOS	Noble inert gases take part in only a few of these.	_______
_______	ACIIC	The oxides of non-metal elements are often this.	_______
_______	UNIVRSAL	This indicator can show the pH of a solution.	_______
_______	CASSIFICATION	The periodic table of elements is a form of this.	_______
_______	GASS	These occur at the right-hand side of the periodic table.	_______
_______	INFLAMMABL	Hydrogen is not suitable for airships because it is this.	_______
_______	ULCANISE	To treat rubber with sulphur to improve its strength.	_______

3. He realised that elements should be arranged in order of their atomic number.

_______	COBINATION	Sodium and chlorine are more commonly found in nature as salt or sodium chloride.	_______
_______	CNDUCT	All metals can do this to heat and electricity.	_______
_______	TRANITION	Another word for the heavy metals between groups II and III of the periodic table.	_______
_______	ON	The number of electrons in the outer shell of an alkali-metal.	_______
_______	ITHIUM	An alkali metal which burns with a red flame.	_______
_______	DÖBREINER	He suggested elements could be grouped in triads (threes)	_______
_______	CATALSTS	Heavy metal elements are often used as these in order to increase the rate of reactions.	_______



ALKALINE EARTH METALS

INTRODUCTION

The metals in group 2A are called alkaline earth metals. This group includes beryllium (Be), magnesium (Mg),calcium (Ca), strontium (Sr), barium (Ba) and radium (Ra). The word “earth” was used for the oxides of group 2A by the ancient scientists. They thought of these oxides as elements. In fact, magnesium, calcium, strontium and barium metals were first isolated from their oxides.

Alkaline earth metals, after the alkali metals, are secondary metals with strong metallic properties. The group 2A elements are less active than those of 1A, whereas they are more active than those of group 3A. Except Be, all form ionic compounds. Be forms mostly covalent compounds. The electrons in their valence shell occupy the s orbitals. Therefore, by losing these electrons through a chemical change they easily gain +2 charge.

In this group, beryllium has similar chemical properties with aluminum, and magnesium has similar chemical properties with lithium.

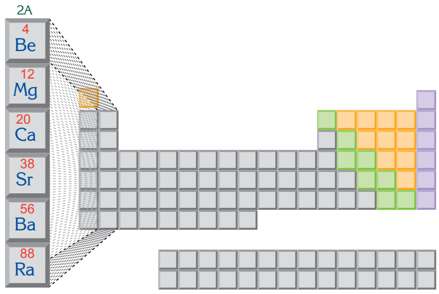


Figure 1 *Alkaline earth metals in the periodic table.*

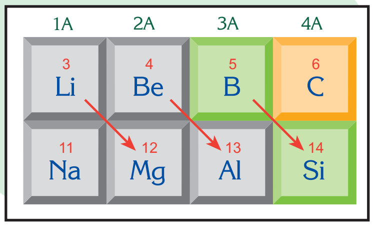


Figure 2 *Diagonal relation between elements in the periodic table.*

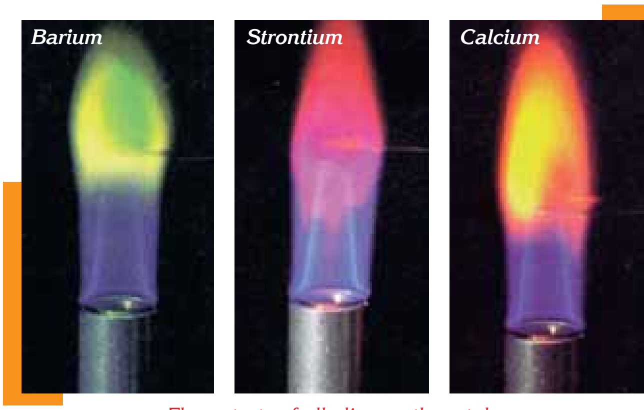
Such cross–wise relationships between elements in the upper side of the periodic table are known as diagonal relations. A similar relationship is seen between boron and silicon as well.

In table 1, some properties of alkaline earth metals are shown.

Name	Beryllium	Magnesium	Calcium
Symbol	Be	Mg	Ca
Atomic number	4	12	20
Atomic mass	9	24,3	40.1
Electron configuration	[He]$2s^2$	[Ne]$3s^2$	[Ar]$4s^2$
Melting point (°C)	1287	650	842
Boiling point (°C)	2469	1090	1484
Density (g/cm^3)	1,85	1.74	1.55
Ist Ionization energy (kj/mol)	899,5	737.7	589.8
Atomic radius (pm)	90	130	174
Common oxidation numbers	+2	+2	+2
Color	lead grey	silvery white	silvery white
Physical state at 25°C	solid	solid	solid
Origin and meaning of name	*beryllos* - beryl	*magnesia* - a name for historical city in Turkey	*calx* - lime
Earth’s crust abundance (%)	$1.9x10^{-4}$	2.9	5

Table 1 *Some properties of alkaline earth metals*

Among the group 2A elements, there are physical relationships which are similar to the group 1A elements. The atomic radius increases and the ionization energy decreases in going from top to bottom. The group 2A elements, except beryllium, do not have electron affinity. The melting point, boiling point and the hardness of the elements decrease from beryllium to barium in the group. Beryllium is the hardest, whereas barium is the softest alkaline earth metal. The members of the group are good conductors of heat. They are malleable and ductile.



Flame tests of alkaline earth metals

Their ions have own characteristic colors in flame tests. Calcium ion has red–orange, strontium ion has dark red and barium ion has light–green color in the flame of a Bunsen burner.

0,2 mol/L metal nitrate	0,2 mol/L sodium salt CO_3^{-2}	0,2 mol/L sodium salt SO_4^{-2}	0,2 mol/L sodium salt CrO_4^{-2}
Mg	White precipitate 12.6×10^{-3}	No precipitate 21.96	No precipitate 119
Ca	White precipitate 1.3×10^{-3}	No precipitate 0.64	No precipitate 13.6
Sr	White precipitate 1.0×10^{-3}	White precipitate 13×10^{-3}	Yellow precipitate 120.4
Ba	White precipitate 1.8×10^{-3}	White precipitate 0.21×10^{-3}	Yellow precipitate 2.53

Table 2 *Solubilities of group 2A salts in 100 g water.*

Compounds with carbonate and oxides of the alkaline earth metals are insoluble in water (Table 2). Excess Mg^{+2} and Ca^{+2} in water means that it is hard. When such hard water is boiled, some amount of precipitate is formed at the bottom of the container. Hard water prevents formation of lather because Ca^{+2} ions and soap form a compound (calcium stearite) which is insoluble in water.

$$\underset{\text{soap}}{2C_{17}H_{35}COONa} + CaSO_4 \rightarrow \underset{\text{calcium stearite}}{(C_{17}H_{35}COO)_2Ca} + Na_2SO_4$$

EXTRA

How are colored fireworks made?

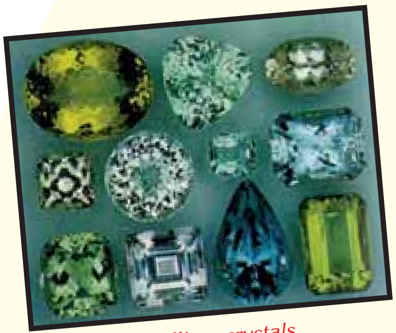
Fireworks existed in ancient China in the ninth century where saltpeter (potassium nitrate), sulfur, and charcoal were mixed to produce dazzling effects. Magnesium burns with a brilliant white light and is widely used in making flares and fireworks. Various other colors can be produced by adding other substances to the flame. Strontium compounds color the flame scarlet and barium compounds produce a yellowish-green color; borax produces a green color, and lithium a purple color.

1. OCCURRENCE

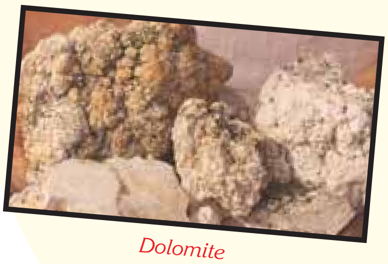
Since the group 2A elements are relatively active metals, they occur in compounds in nature.

Beryllium, Be

Beryllium occurs in trace amount in nature. The most important ore of beryllium is beryl, $Be_3Al_2(SiO_3)_6$.



Beryllium crystals



Dolomite

Magnesium, Mg

The principal useful ores o f magnesium are dolomite ($CaCO_3 \cdot MgCO_3$ a double salt), carnallite, ($KCl \cdot MgCl_2 \cdot 6H_2O$) and epsom salt ($MgSO_4 \cdot 7H_2O$) which is found in mineral water. Epsom salts are used as a purgative, as a dressing for cotton goods and in dye.

The magnesium metal is the center of the structure of chlorophyll.

Calcium, Ca

Calcium compounds are widely distributed in nature, occurring as limestone or marble ($CaCO_3$), gypsum ($CaSO_4 \cdot 2H_2O$) and fluorite (CaF_2). Salts of sulfate, silicate and phosphate are also found in the earth's crust.

Strontium, Barium and Radium; Sr, Ba, Ra

They are found in trace amount in nature. Strontium occurs as celestite ($SrSO_4$) and strontianite ($SrCO_3$), barium occurs as barytes ($BaSO_4$) and witherite ($BaCO_3$). All the isotopes of radium are radioactive. The isotope ^{226}Ra is the most stable isotope with a 1600-year half–life. The radium element is a side product of the natural decay of ^{238}U.

Preparation of Calcium

Preparation of calcium can be given as a good example of preparations of group 2A elements.

The most important preparation method of calcium is the electrolysis of a molten mixture of $CaCl_2$–CaF_2 · CaF_2 is used to reduce the melting point of $CaCl_2$ below the melting point of calcium metal (840°C). The cathode is an iron rod which just dips below the surface of the calcium chloride. As the calcium forms on the iron rod, the iron rod is withdrawn and a stick of calcium is gradually formed. Chlorine gas is evolved at anode. The anode is made of graphite because chlorine gas attacks most metals. The gas is a valuable by–product.

The other method is reduction of calcium oxide with aluminum. Calcium metal is used mainly in some alloys.

$$2Al(s) + 4CaO(s) \xrightarrow{1200°C} CaO \,.\, Al_2O_3(s) + 3Ca(s)$$

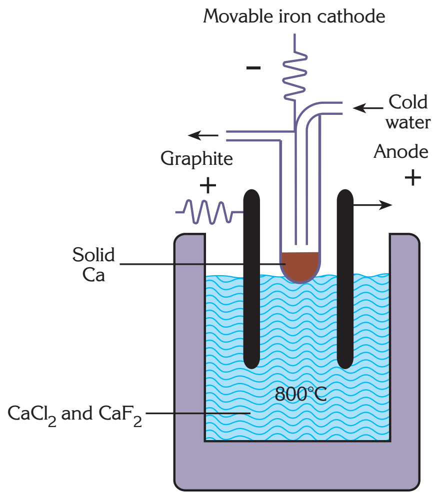


Figure 3 Manufacture of calcium metal diagram

2. CHEMICAL PROPERTIES

All the members of the group are covered by their oxides when they are exposed to air. The alkaline earth metals burn in air with bright flame. As a result of burning in air, their oxides and nitrides are formed. The compounds of all the alkaline earth metals are similar in composition, they all form oxides (MO), hydroxides ($M(OH)_2$), carbonates (MCO_3), sulfates (MSO_4), and other compounds (M is any alkaline earth metal). The alkaline earth metals are strong oxidizing agents.

Activity of the metals increases from Be to Ba. The elements Ca, Sr and Ba undergo reaction with water easily, whereas the reaction of Be and Mg with water is difficult. In order to increase the activity of Mg metal, the Mg metal is transformed into an amalgam. Thus, since the oxide of Mg is not formed on the surface of the metal, its activity increases in the chemical reactions.

Reactions

1. All alkaline earth metals, except beryllium, react with H_2 gas in hot medium to produce hydrides, like alkali metals.

   $$M(s) + H_2(g) \xrightarrow{heat} MH_2(s)$$
   (M : alkaline earth metals)

   $$Mg(s) + H_2(g) \xrightarrow{heat} MgH_2(s)$$

   $$Ca(s) + H_2(g) \xrightarrow{heat} CaH_2(s)$$

   Only CaH_2, of all hydrides formed, reacts with water by releasing heat.

   $$CaH_2(s) + 2H_2O(l) \longrightarrow Ca(OH)_2(s) + 2H_2(g)$$
   (Figure 4)

Figure 4 CaH_2 is the only alkaline earth metal hydride which reacts with water.

2. Ca, Sr and Ba react with water, like alkali metals, at room temperature to produce metal hydroxides and hydrogen gas.

$M(s) + 2H_2O(l) \rightarrow M^{+2}(aq) + 2OH^-(aq) + H_2(g)$
(M : alkaline earth metals)

$Ca(s) + 2H_2O(l) \rightarrow Ca^{+2}(aq) + 2OH^-(aq) + H_2(g)$

Magnesium metal reacts slowly with boiling water. The reaction of beryllium with water is very difficult.

$Mg(s) + 2H_2O(l) \longrightarrow Mg(OH)_2(s) + H_2(g)$

$Be(s) + 2H_2O(l) \longrightarrow Be(OH)_2(s) + H_2(g)$

All formed hydroxides of this group lose water to form their anhydrates (oxides) in thermal decomposition.

$Ca(OH)_2 \xrightarrow{heat} CaO + H_2O$

3. They form oxides as a result of their reactions with oxygen, in MO formula (Figure 5).

$2M(s) + O_2(g) \longrightarrow 2MO(s)$

$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$

$2Sr(s) + O_2(g) \longrightarrow 2SrO(s)$

These oxides are slightly soluble in water and solutions of these oxides, except BeO, show basic properties.

Peroxides of Ca, Sr and Ba are produced by heating their oxides in oxygen atmosphere. Barium peroxide can be obtained easily by the following reaction

$BaO + 1/2O_2 \xrightarrow{500–600°C} BaO_2$

If the burning process occurs in air, nitrides, hydroxides and carbonates are formed, besides oxides.

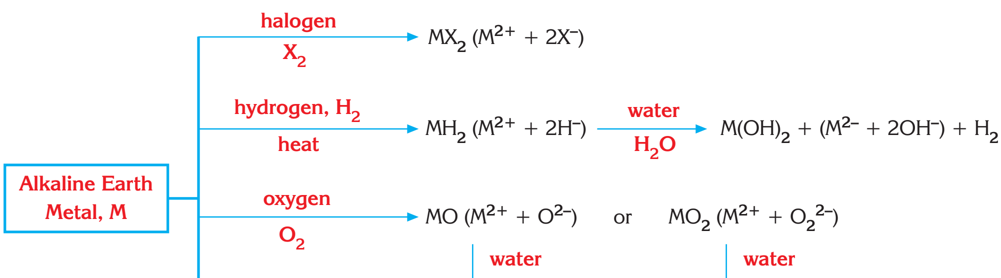


Figure 5
Summary of reactions of Alkaline Earth Metals.

Burning magnesium

4. All alkaline earth metals give direct reactions with halogens to produce metal halides.

$M(s) + X_2(g) \longrightarrow MX_2(s)$

$Be(s) + Cl_2(g) \longrightarrow BeCl_2(s)$

$Mg(s) + Cl_2(g) \longrightarrow MgCl_2(s)$

5. The reactions of the group 2A elements with acids like HCl and H_2SO_4, produce salts and H_2 gas.

$Ca(s) + 2HCl(aq) \longrightarrow CaCl_2(s) + H_2(g)$

$Ba(s) + H_2SO_4(aq) \longrightarrow BaSO_4(s) + H_2(g)$

While magnesium reacts with dilute H_2SO_4 by giving H_2 gas, it reacts with hot and concentrated H_2SO_4 by producing SO_2 gas.

$Mg(s) + H_2SO_4(dil.) \rightarrow MgSO_4(s) + H_2(g)$

$Mg(s) + 2H_2SO_4(aq)(conc.) \rightarrow MgSO_4(aq) + SO_2(g) + 2H_2O(l)$

Similarly, the reaction of magnesium metal with 1% HNO_3 produces H_2 gas, but the reaction of magnesium with concentrated HNO_3(60%) produces NO_2 gas.

$Mg(s) + 2HNO_3(1\%) \rightarrow Mg(NO_3)_2(aq) + H_2(g)$

$Mg(s) + 4HNO_3(60\%) \rightarrow Mg(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$

On the other hand, under normal conditions beryllium does not react with HNO_3 because of the layer of its oxide on the surface of the metal.

6. All oxides and hydroxides of alkaline earth metals, except beryllium, show basic properties whereas beryllium metal, its oxide and its hydroxide show amphoteric properties. In other word, Be, BeO and $Be(OH)_2$ show either acidic or basic properties.

$Be(s) + 2HCl(aq) \longrightarrow BeCl_2 + H_2(g)$

$Be(s) + 2NaOH(aq) + 2H_2O(l) \longrightarrow Na_2[Be(OH)_4](aq) + H_2(g)$

3. COMPOUNDS

The compounds of alkaline earth metals are ionic. They have high melting and boiling points. They are thermally stable. Their salts are less soluble in water than those of alkali metals. And the solubilities of salts increase from beryllium to barium.

Most of the compounds of beryllium and some compounds of magnesium show a tendency to covalency, so they differ from other compounds in the group.

Thermally stable means that the compound can't be decomposed at the temperature of a normal Bunsen flame (approximately 1300K).

Now, the compounds of alkaline earth metals will be examined with the help of the most important calcium compounds.

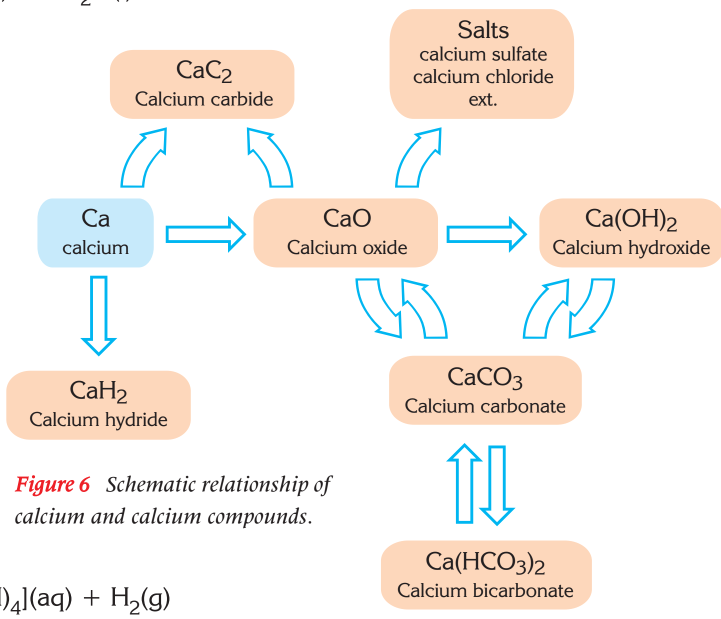


Figure 6 *Schematic relationship of calcium and calcium compounds.*

Group 2A

Carbonates (CO_3^{2-})

Mg $MgCO_3(s) \longrightarrow MgO(s) + CO_2(g)$

↓ ✔ *Successively higher temperatures needed to decompose*

Ba

Hydroxides (OH^-)

Mg $Mg(OH)_2(s) \longrightarrow MgO(s) + H_2O(g)$

↓

Ba ↓ ✔ *More stable*

Hydrogencarbonates (HCO_3^-)

✔ *Too unstable to exist as solids – exist only in solution.*

Nitrates (NO_3^-)

$2Mg(NO_3)_2(s) \longrightarrow 2MgO(s) + 4NO_2(g) + O_2(g)$

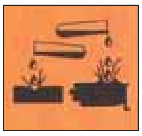
↓ ✔ *More stable*

Sulfates (SO_4^{2-}) and halides (X^-)

✔ *All thermally stable*

Table 3 Effect of heat on compounds of group 2A elements

1. Calcium oxide, CaO and calcium hydroxide, $Ca(OH)_2$



Calcium oxide (lime) is used in the production of cement, and is obtained by heating calcium carbonate (limestone) at 900°C.

$$CaCO_3(s) \xrightarrow{900°C} CaO(s) + CO_2(g)$$

Addition of calcium oxide to water is called slaking, and the product, $Ca(OH)_2$, is called slaked lime (lime water).

$$CaO + H_2O \longrightarrow Ca(OH)_2 + 16kcal$$

$Ca(OH)_2$ is used in the preparation of mortar (slaked lime plus sand) which sets to a solid by reconversion of the $Ca(OH)_2$ to $CaCO_3$ as a result of the reaction with CO_2 in air.

$$Ca(OH)_2(s) + CO_2(g) \longrightarrow CaCO_3(s) + H_2O(g)$$

Stalactites and stalagmites

2. Calcium carbonate, $CaCO_3$

It occurs naturally as marble, limestone, chalk and calcite. As mentioned above, it forms by evolving CO_2 gas from $Ca(OH)_2$ solution.

Underground water containing CO_2 dissolve some amount of $CaCO_3$ and the equilibrium below is established.

$$CaCO_3(s) + CO_2(g) + H_2O(l) \rightleftharpoons Ca^{+2}(aq) + 2HCO_3^-(aq)$$

When such underground water reaches the earth's surface, the pressure on it decreases. Because of the decrease in solubility of CO_2 in water, the equilibrium reaction shifts to the left, that is, $CaCO_3$ precipitate is formed. Formation of stalactites and stalagmites is a good example of this process.

Pamukkale

3. Calcium carbide (CaC_2)

One of the calcium compounds is calcium carbide, CaC_2, which is found in the form of small gray solid pieces, used widely in industry. Calcium carbide is prepared commercially by the reaction of calcium oxide (lime) and coke at high temperature.

$$CaO(s) + 3C(s) \longrightarrow CaC_2(s) + CO(g)$$

The reaction of calcium carbide with water produces acetylene, C_2H_2, which is used in welding.

$$CaC_2(s) + 2H_2O(l) \longrightarrow C_2H_2(g) + Ca(OH)_2(s)$$

Calcium carbide mineral

4. Calcium sulfate, $CaSO_4$

In nature, its anhydrate ($CaSO_4$) and dihydrate forms ($CaSO_4 . 2H_2O$) are called gypsum and alabaster.

Gypsum ($CaSO_4 . 2H_2O$) is a white substance which is used commercially for fabrication into wallboard. When gypsum is heated above 100°C, it loses 3/4 of its water of crystallization, forming the powdered substance $CaSO_4 . 1/2H_2O$, known as plaster of Paris.

$$\underset{\text{gypsum}}{2CaSO_4{\cdot}2H_2O} \xrightarrow{120\text{–}150^\circ C} \underset{\text{plaster of paris}}{2CaSO_4{\cdot}1/2H_2O} + 3H_2O$$

By mixing with water, the small crystals of plaster of Paris dissolve and then re-crystallize as long needles of $CaSO_4 . 2H_2O$

$$CaSO_4{\cdot}1/2H_2O + 3/2H_2O \xrightarrow{20^\circ C} CaSO_4{\cdot}2H_2O$$

Calcium sulfate, $CaSO_4$ is a constituent of cement. Additionally, it is used in decoration as plaster of Paris, and in dentistry to mold teeth.

Formula	Common name	Uses
MgO		In lining of oven to manufacture rubber and dye
$Mg(OH)_2$	Milk of magnesia	In medicine as anti-acid
$MgSO_4{\cdot}7H_2O$	Epsom salt	In manufacture of dye
$CaSO_4{\cdot}2H_2O$	Gypsum	In buildings
$CaSO_4{\cdot}1/2H_2O$	Plaster of Paris	In manufacture of gypsum
CaO	Lime	In manufacture of cement and casting
CaC_2	Carbide	In welding
$CaCO_3$	Limestone	In buildings
$Ca(OH)_2$	Limewater	In plasters
BaO_2		In manufacture of paper and dye
$BaSO_4$	Baryte	In X-ray photography
$Sr(NO_3)_2$		In fireworks as red color

Table 4 Uses of some compounds of alkaline earth metals

USES

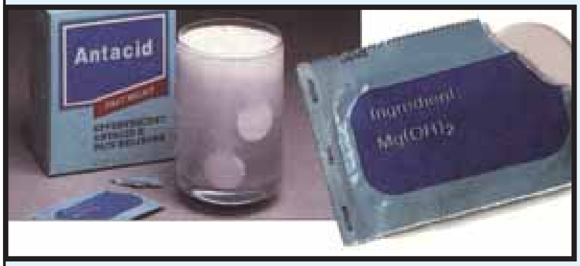


Figure 7 *$Mg(OH)_2$, milk of magnesia is used as a stomach antacid.*

Metallic beryllium is used to prepare some special alloys. About 2% beryllium in copper produces a hard alloy especially suited for use in springs. The metal is used in making glass for X-ray tubes, so x-rays readily penetrate elements with low atomic number. Beryllium has the mechanical properties of a very light element.

$BaSO_4$ is used in medicine in X-ray studies of the intestines.

Since magnesium is a light metal, it is used in the production of planes, missiles and some light household items. In alloys, it is used to increase the hardness, durability and resistance to corrosion. The light produced by the burning of magnesium tape with oxygen, is used as a flash in photography. Mg is also used as an additive in rocket fuels and signal rockets. $Mg(OH)_2$, milk of magnesia is used as a stomach antacid.

The compounds of calcium, hydroxides, carbonates and oxides, are widely used as construction materials and as bleaching agents. Calcium oxide is the main compound in the production of some chemicals, steel, glass, paper and sugar. Calcium carbide (CaC_2) is used to produce acetylene gas for welding.

Acetylene gas produced from calcium carbide.

Strontium has many important applications in industry. The salts of strontium are used in signal rockets and fireworks to produce red light.

Barium is used as a gas absorbent in vacuum tubes. Since the alloys of barium with nickel donate electrons by heating they are used in vacuum tubes and in ignitors. Barium nitrate ($Ba(NO_3)_2$) and barium chlorate ($Ba(ClO_3)_2$) are used to obtain green light in fireworks. $BaSO_4$ (barium meal) is used to take the photos of stomach and intestines because, like all elements with large atomic number, it does not allow X-rays to pass. In fact, the Ba^{2+} ion is rather toxic, but since the solubility of $BaSO_4$ in water is very low, its poisonous effect disappears. Radium is used in selfluminous paints and in radiotherapy and as a neutron source.

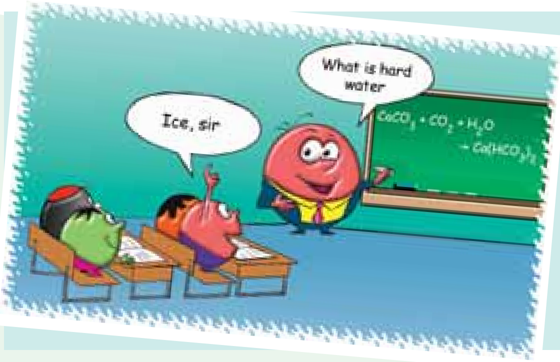
Alloy of magnesium is used in the manufacture of airplanes.

The main component of cement is CaO.

reading

Water Softening

Water containing appreciable amounts of Ca^{2+} and Mg^{2+} is called hard water. The anions that are usually present with these cations are Cl^-, SO_4^{2-}, and HCO_3^-. Water containing only very small concentrations of Ca^{2+} and Mg^{2+} is called soft water.



Most natural water, particularly in limestone regions, is hard. The use of hard water for domestic purposes and in industry presents several problems. Since the substances which cause water hardness must be removed before the water is used. The removal of the metallic ions responsible for the hardness is called water-softening.

Water softening methods depend on the type of hardness. Hard waters are classified as temporary-hard water and permanently hard water.

Temporary hard water: CO_2 dissolved in water or acid rain water dissolves the limestone ($CaCO_3$) and dolomite ($MgCO_3 \cdot CaCO_3$) in the earth's crust Ca^{2+} and Mg^{2+} ions pass into water as soluble hydrogen carbonate salts $Ca(HCO_3)_2$ and $Mg(HCO_3)_2$.

$$MgCO_3 \cdot CaCO_3(s) + CO_2(g) + H_2O(l) \longrightarrow Mg(HCO_3)_2(aq) + Ca(HCO_3)_2(aq)$$

This kind of water is known as temporary hard water and it can be softened by boiling a few minutes. When hard water is heated, Ca^{2+} ions precipitate as $CaCO_3$.

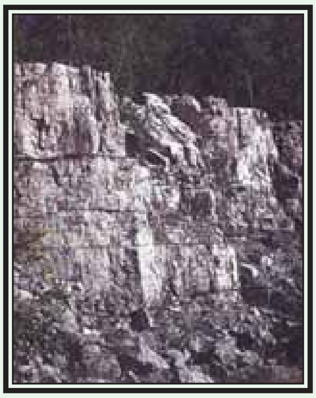
$$Ca^{2+}(aq) + 2HCO_3^-(aq) \rightarrow CaCO_3(s) + CO_2(aq) + H_2O(l)$$

The other method is to add a base solution.

$$Ca(HCO_3)_2(aq) + Ca(OH)_2(aq) \rightarrow 2CaCO_3(s) + 2H_2O(l)$$

Permanently hard water: Dissolved $CaCl_2$ and $CaSO_4$ salts can't be removed from water by boiling. This water is called permanently hard water. Permenant hardness can be removed by adding washing soda, Na_2CO_3. In this method, Ca^{2+} and Mg^{2+} ions are precipitated as their carbonates.

$$CaSO_4(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + Na_2SO_4(aq)$$



For large-scale treatment of water, the ion exchange method is used. The diagram of the system is shown in the figure given. When hard water is taken into this apparatus, on the way down the Ca^{2+} ions are swapped for sodium ions. The Ca^{2+} ions get stuck in the zeolite ($NaAlSi_2O_6$). Na^+ ions don't cause hardness, but come out in the water at the bottom.

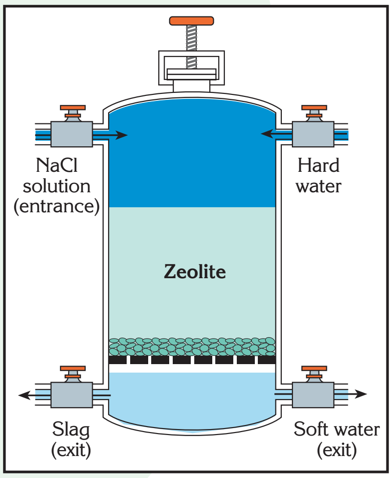


Diagram of softening of hard water by changing of ion method in industry.

Over time, the amount of Ca^{2+} in zeolite will increase. To remove more Ca^{2+} ions, NaCl solution is removed at the tap and the process above is reversed. In this way, sodium zeolite is regenerated and is ready for use again.

The ion-exchange method of water softening is effective for both kinds of water (temporarily hard and permanently hard water).

Disadvantages of hard water	Advantages of hard water
Difficult to form lather with soap.	Some people prefer the taste.
Scum forms in a reaction which wastes soap.	Calcium in the water is good for children's teeth and bones.
Scale (a hard crust) forms inside kettles. This wastes energy when you boil your kettle.	Helps to reduce heart disease.
Hot water pipes "fur up" on the inside. The scale formed can even block up pipes completely.	

SUPPLEMENTARY QUESTIONS

1. Write the electron configuration of $_4Be$.

2. Write the elements in the alkaline earth metals group.

3. What is diagonal relation in the periodic table?

4. Which compound of alkaline earth metals is used as a purgative?

5. Which element is the hardest in the alkaline earth metals group?

6. What is the meaning of hard water? What are the negative effects of hard water?

7. What mass of CaH_2 is produced at the end of the reaction of 5 g of Ca with excess H_2?

8. What mass of Mg reacts with excess water to produce 8 g of MgO?

9. What mass of Sr reacts with excess O_2 to produce 10.45 g of SrO?

10. What mass of $BeCl_2$ is produced at the end of the reaction of 1.8 g of Be with excess Cl_2 with 80% efficiency?

11. What volume of H_2 gas is produced at the end of the reaction of 1.2 g of Mg with 98 g of H_2SO_4 49% by mass at STP?

12. What volume of H_2 gas is produced at the end of the reaction of 72 g of Be with 73 g of HCl at STP?

13. Give equations for all possible reactions of the oxidation of magnesium in air.

14. Which member of group 2A is radioactive?

15. Complete the following equations

a. $CaC_2 + H_2O \longrightarrow$

b. $CaCO_3(s) \xrightarrow{900\ °C}$

c. $BaO_2 \xrightarrow{heat}$

d. $CaO + C \longrightarrow$

e. $Mg(OH)_2 + CO_2 \longrightarrow$

f. $CaH_2 + H_2O \longrightarrow$

g. $MgO + H_2O_2 \longrightarrow$

16. Complete the following equations.

a. $Ca + Cl_2 \longrightarrow$

b. $Ca + H_2SO_4 \longrightarrow$

c. $Ca + O_2 \longrightarrow$

d. $Ca + H_2O \longrightarrow$

e. $Mg + HNO_3(conc) \longrightarrow$

f. $Be + HCl \longrightarrow$

17. How can we explain the different chemical tendencies of beryllium?

18. What is the mass percentage of beryllium in ore that contains 42% beryl minerals?

19. Why is CaF_2 added to $CaCl_2$ solution in the production of Ca by the electrolysis method?

20. Find the alkaline earth metal that has given the properties.

a. All its electrons are found in s orbitals.

b. It is radioactive.

c. Its properties are similar to those of aluminum.

d. It is found in chlorophyll.

e. It is very inert to water.

f. Its compounds are widely used in building.

g. It causes water hardness and doesn't react with cold H_2SO_4.

21. The α-particles emitted in the radioactive decay of radium -226 can be counted with of a Geiger counter.

a. Write the equation of this event adding all the mass numbers, atomic numbers and symbols.

b. Each α-particle gains electrons to form helium gas. It is found that 1.82×10^{12} α-particles give 6.75×10^{-3} cm^3 of helium at S.T.P. By using these data, obtain a value for Avogadro's constant.

22. The solubility of barium hydroxide, $Ba(OH)_2$, at 25°C is 0.24 g/L.

a. Calculate the molar concentration of the saturated aqueous solution.

b. If it is assumed that the solute is completely ionized, calculate the hydroxide ion concentration of the solution.

23. Bottles containing aqueous barium hydroxide need to be kept firmly stoppered or a white deposit forms on the surface.

a. What is this white deposit?

b. Explain, with an equation, how it is formed?

24. At 80°C, by using 100 g H_2O and 60 g magnesium sulfate monohydrate, a saturated solution was prepared. If we cool this solution to 20°C, how many g of magnesium sulfate heptahydrate precipitate? The solubility of $MgCl_2$ is 35.1 g/100mL H_2O at 20°C.

25. What is the mass of $CaCl_2 \cdot 6H_2O$ we need to add to 47 mL of 25% Na_2CO_3 solution (d = 1,08 g/mL) in order to get a solution which contains 10% Na_2CO_3 by mass?

26. Write the chemical equations for each step in following reaction schemas.

a. $Mg \longrightarrow MgO \longrightarrow MgSO_4$

b. $Ca \longrightarrow CaH_2 \longrightarrow Ca(OH)_2 \longrightarrow Ca(NO_3)_2 \longrightarrow CaCO_3$

MULTIPLE CHOICE QUESTIONS

1. What is the common oxidation number of the alkaline earth metals?

A) –2 B) +2 C) +1 D) –1 E) +3

2. Which is the most abundant in the earth's crust?

A) Be B) Mg C) Ca D) Sr E) Ba

3. Which one has the greatest ionization energy?

A) Be B) Mg C) Ca D) Sr E) Ba

4. What is the flame test color of barium element?

A) White B) Red C) Orange D) Light blue E) Light green

5. Which of the following compounds show amphoteric properties?

A) BaO B) MgO C) Sr D) BeO E) CaO

6. Which compound is used to produce acetylene gas for welding?

A) CaC_2 B) BaC_2 C) MgC_2 D) Al_4C_3 E) CO_2

7. Which of the following solutions is used in X-ray photography?

A) $CaSO_4$ B) $BaSO_4$ C) $Ba(NO_3)_2$ D) SrO E) CaO

8. Which color is observed when Mg is burnt?

A) Yellow B) Blue C) Red D) Gray E) White

9. Which gas is produced when an alkaline earth metal reacts with water?

A) Oxygen B) Hydrogen C) Carbon dioxide D) Nitrogen E) Chlorine

10. Which of the following does not react easily?

A) Be B) Mg C) Ca D) Ba E) Sr

11. Which gas is produced when calcium reacts with sulfuric acid?

A) Hydrogen B) Hydrogen sulfide C) Sulfur dioxide D) Oxygen E) Sulfur trioxide

12. Which is not a property of alkaline earth metals?

A) They react with water
B) They react with acids to produce H_2 gas
C) They do not conduct heat and electricity.
D) They are active metals.
E) They are in group 2A of the periodic table.

13. Which member of group 2A is the least active metal?

A) Barium B) Strontium C) Beryllium D) Calcium E) Sodium

14. Which one is correct about alkaline earth metals?

A) They produce oxygen gas with water.
B) They do not conduct electricity.
C) They form salts with halogens.
D) They have +1 oxidation number in their compounds.
E) They are more active than the metals in group1A.

15. Which of the following activity orders is correct?

A) Mg > Ba > Ca B) Be > Ca > Mg C) Ba > Ca > Mg D) Sr > Ba > Ca E) Ca > Mg > Sr

16. Which one is used in fireworks?

A) Hydrogen B) Gold C) Magnesium D) Oxygen E) Silver

17. What is the formula of limestone?

A) CaO B) $CaCO_3$ C) $Ca(OH)_2$ D) $CaSO_4.2H_2O$ E) CaI_2

18. Which one is the electron configuration of ${}_{20}Ca^{2+}$?

A) $1s^22s^22p^6$ B) $1s^2s2p^63s^23p^64s^2$ C) $1s^2s^22p^2$ D) $1s^22s^22p^63s^23p^6$ E) $1s^22s^22p^63s^23p^4$

19. I. Cl_2
II. H_2O
III. Na_2O
Which of the above react(s) with calcium at normal conditions?

A) I only B) III only C) I and II D) II and III E) I, II and III

20. What is the sum of the electrons in n = 2 and n = 3 (2nd and 3rd energy levels) in barium atom? (Ba_{56})

A) 36 B) 14 C) 18 D) 8 E) 26

21.

	Common name	Formula
I.	Limewater	$Ca(OH)_2$
II.	Beryl	$Be_3Al_2(SiO_3)_6$
III.	Gypsum of Paris	$CaSO_4 . 1/2H_2O$

Which of the above matches is/are true?

A) I only B) III only C) I and II D) II and III E) I, II and III

22. Which ion does not cause water hardness?

A) Ca^{2+} B) CO_3^{2-} C) HCO_3^- D) Mg^{2+} E) Na^+

23.
I. $CaH_2 + H_2O \longrightarrow$
II. $Be + NaOH \longrightarrow$
III. $Mg + HNO_3(conc) \longrightarrow$

Which give(s) H_2 gas?

A) I only B) III only C) I and II D) II and III E) I, II and III

24. Which example is not correct for the given reaction type?

A) $CaSO_4 . 2H_2O \longrightarrow CaSO_4 . 1/2H_2O + 3/2H_2O$ Dehydration

B) $CaCO_3 \longrightarrow CaO + CO_2$ Decomposition

C) $CaSO_4 + 2C_{17}H_{35}COONa \longrightarrow (C_{17}H_{35}COO)_2Ca + Na_2SO_4$ Precipitation

D) $Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$ Redox

E) $2BaO + O_2 \longrightarrow 2BaO_2$ Synthesis

25.

	Ion	Flame Test
I.	Calcium	Blue
II.	Strontium	Dark red
III.	Barium	Light green

Which flame test color is given wrong?

A) I only B) I and II C) I and III D) II and III E) I, II and III

26. Which one of the following equations represents the reaction that occurs when calcium nitrate is heated strongly?

A) $Ca(NO_3)_2 \longrightarrow Ca(NO_2)_2 + O_2$

B) $Ca(NO_3)_2 \longrightarrow CaO + N_2O + 2O_2$

C) $Ca(NO_3)_2 \longrightarrow CaO_2 + 2NO_2$

D) $2Ca(NO_3)_2 \longrightarrow 2CaO + 4NO_2 + O_2$

E) $3Ca(NO_3)_2 \longrightarrow Ca_3N_2 + 4NO_2 + O_2$

27. In hospital, barium sulfate is used in taking x-ray photographs of the intestines. Why is the sulfate used rather than other compounds of barium?

I. Soluble barium compounds are poisonous
II. Barium sulfate reacts with organic material in the body
III. Barium sulfate forms sulphuric acid with the acid in the stomach.

A) I only B) I and II C) II and III D) I and III E) I, II and III

28. What conclusions can be drawn from the observation that $BaSO_4$ is precipitated immediately when solutions of barium chloride and sodium sulfate are mixed?

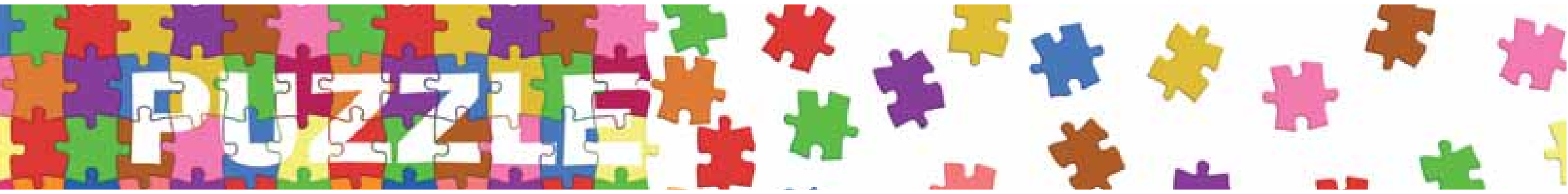
A) $BaSO_4$ is a predominantly covalent molecule.

B) $BaSO_4$ is strongly hydrated.

C) $BaSO_4$ is very soluble in H_2O.

D) Free $Ba^{2+}(aq)$ and $SO_4^{2-}(aq)$ ions probably exists in the initial solutions.

E) $Ba(OH)_2$ is a very weak base.



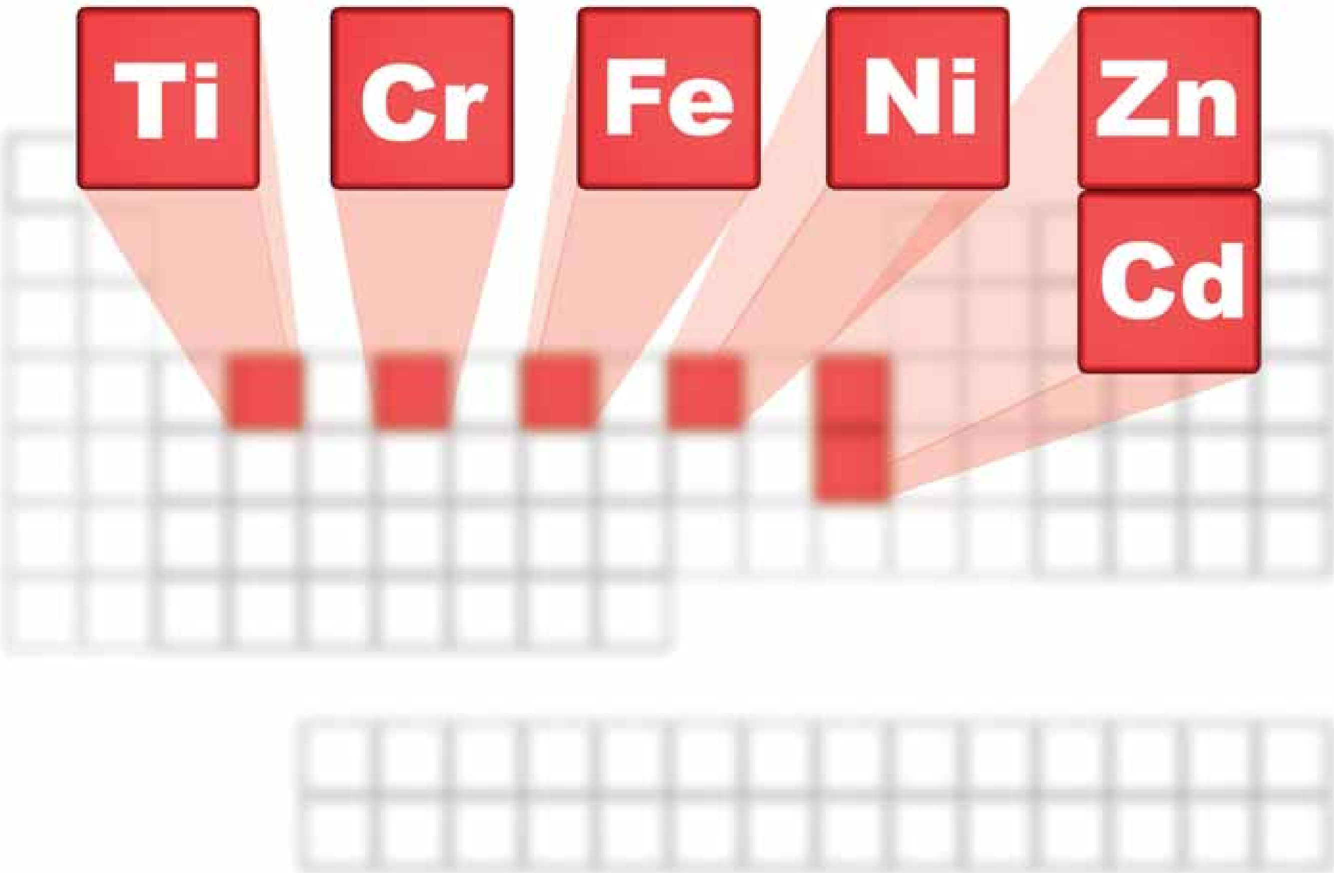
The definitions of the following words have become confused. Can you sort them out? Put the pairs of letters and numbers in the grid provided.

1.	MAGNESIUM	A	A balanced equation shows this for each reactant and product.
2.	HETEROGENOUS	B	A type of reversible reaction in which a compound is divided into other compounds or element.
3.	SYMBOLS	C	This always accompanies an endothermic reaction.
4.	COMPLETION	D	An inhibitor is this type of catalyst.
5.	FORMULA	E	The first member of group 2A in the periodic table.
6.	WORD EQUATION	F	One of the elements affect the hardness of water.
7.	MOLES	G	Reactions are this if the products can turn back into the reactants again.
8.	CALCIUM	H	These are the new substances formed during a reaction.
9.	TEMPERATURE RISE	I	This type of catalyst is in a different physical state from the reactants.
10.	DISSOCIATION	J	This measurement of reactants and products is always identical as long as all reactants and products are taken into account.
11.	PRODUCTS	K	A balanced equation shows the number of these that are reacting and the number of these that are being produced.
12.	COLLISION	L	A type of reaction where two or more simple substances join together to form a single product.
13.	CLOSED	M	This type of reaction has such a fast rate that it occurs almost instantaneously.
14.	EQUILIBRIUM	N	We can describe what happens in a chemical reaction by writing this.
15.	RATE	O	When the end point of a titration is reached the acid has done this to the alkali or vice versa.
16.	TEMPERATURE FALL	P	A system in which no chemicals can escape or enter.
17.	MASS	Q	This type of reaction involves a substance splitting up when it reacts with water.
18.	SYNTHESIS	R	These are used as a shorthand method of writing something, i.e. the formula of a compound might be written H_2O
19.	NEGATIVE	S	Endothermic reactions involve this type of heat change.
20.	CONDENSATION	T	This will always accompany an exothermic reaction.
21.	REVERSIBLE	U	Reactions proceed at a fast rate when the reactants are increased in this way.
22.	CONCENTRATION	V	To go to this a chemical reaction must continue until one or all of the reactants are used up and their products must not react together.
23.	NEUTRALISED	W	These lower the energy barrier that stops reactions happening which makes them happen more quickly and easily.
24.	SURFACE AREA	X	Increasing the temperature by 10 degrees Celsius doubles this for many reactions.
25.	HYDROLYSIS	Y	An element the most important mineral is, Beryl.
26.	VANADIUM	Z	This theory explains why altering the conditions under which a reaction takes place affects its rate.
27.	POSITIVE	AA	The starting materials in a chemical reaction.
28.	EXPLOSION	BB	Increasing this will increase the rate of a reaction.
29.	REACTANTS	CC	A stage reached in a reversible reaction in a closed system when the forward and backward reactions take place at the same rate.
30.	CATALYSTS	DD	This type of reaction is the reverse of hydrolysis.

A	B	C	D	E	F	G	H	I	J	K	L	M	N	O

P	Q	R	S	T	U	V	W	X	Y	Z	AA	BB	CC	DD

NOTES



SOME TRANSITION METALS

INTRODUCTION

The transition metals lie between groups 2A and 3A of the periodic table. These elements are often called block elements because their electronic configuration ends with d orbitals.

Transition metals are typical metals. Like other metals they are hard, dense and shiny. They are malleable and ductile. They are also good conductors of heat and electricity. However they have some differences from group 1A and 2A metals. Transition metals are less reactive than group 1A and 2A elements. They mosty form colored compounds whereas most compounds of group 1A and 2A metals are white.

Melting and boiling points of transition metals are higher than those of group 1A and 2A metals.

Oxidation states of group 1A and 2A metals are constant, +1 and +2 respectively. On the other hand, transition metals have more than one oxidation state. For example, chromium can take any charge from +1 to +6.

Now, let's study some of the transition metals in detail.

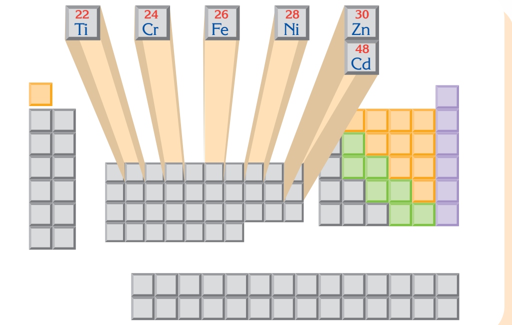


Figure 1 *Some transition metals*

1. IRON

Iron was used in Egypt as early as 3000 B.C. Pure iron is a silvery white colored, lustrous, soft metal with important magnetic properties. It is malleable and ductile. The pure metal is very reactive chemically, and rapidly corrodes, especially in moist air or at high temperatures.

Its density is 7.87 g/cm^3, melting point is 1538°C and boiling point is 2861°C.

Name	Iron	Nickel	Chromium	Titanium	Zinc	Cadmium
Symbol	Fe	Ni	Cr	Ti	Zn	Cd
Atomic number	26	28	24	22	30	48
Atomic mass	55.9	58.7	52	47.9	65.4	112.4
Electron configuration	$[Ar]3d^6 4s^2$	$[Ar]3d^8 4s^2$	$[Ar]4s^1 3d^5$	$[Ar]3d^2 4s^2$	$[Ar]3d^{10} 4s^2$	$[Kr]4d^{10} 5s^2$
Melting point (°C)	1538	1455	1907	1668	419.5	321.1
Boiling point (°C)	2861	2913	2671	3287	907	767
Density (g/cm^3)	7.87	8.9	7.19	4.5	7.14	8.65
Ist Ionization energy (kj/mol)	762.5	737.1	652.9	658.8	906.4	867.8
Atomic radius (pm)	125	121	127	136	131	148
Common oxidation numbers	+2, +3	+2	+2, +3, +6	+2, +3, +4	+2	+2
Color	greyish-tinge	silvery-tinge	silvery metallic	silvery metallic	bluish pale grey	silvery grey
Physical state at 25°C	solid	solid	solid	solid	solid	solid
Origin and meaning of name	*ferrum* - iron	*kupfer nickel* - Devil's copper	*chroma* - color	the sons of the Earth in mythology	German word zink	*cadmia* - calamine
Earth's crust abundance (%)	6.3	$9x10^{-3}$	$1.4x10^{-2}$	0.66	$7.9x10^{-3}$	$1.5x10^{-5}$

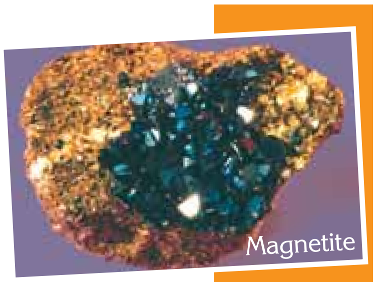
Table 1 *Some transition metals and their properties.*

1.1 OCCURRENCE

Iron is the second most abundant metal (6.2% by mass) in the earth's crust. But iron is not found in elemental form in nature. Elemental iron is found only in meteors with cobalt and nickel as alloy. Iron comprises 4% of volcanic rock and can be found abundantly in the center of the earth.

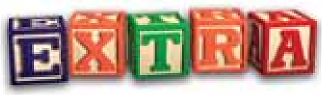
The meteor (weight: 85 kg) that fell in Eşkişehir – Sivrihisar in Turkey.

Iron is found in most clays, sandstones and granites. Hematite (Fe_2O_3), magnetite (Fe_3O_4) pyrite (FeS_2) and siderite ($FeCO_3$) are common ores of iron present in the earth's crust. Hematite often occurs as the reddish-brown hydrate ($Fe_2O_3 \cdot 3H_2O$) called limonite. Brown or red colored rocks take their colors from iron oxides. At the same time, iron is the basic component of hemoglobin, which gives the color of red blood cells.



What is fool's gold?

Pyrite (FeS_2) is a mineral popularly known as "fool's gold." Because of its metallic luster and pale brass yellow color, it is often mistaken for gold. Real gold is much heavier, softer, not brittle, and not grooved.

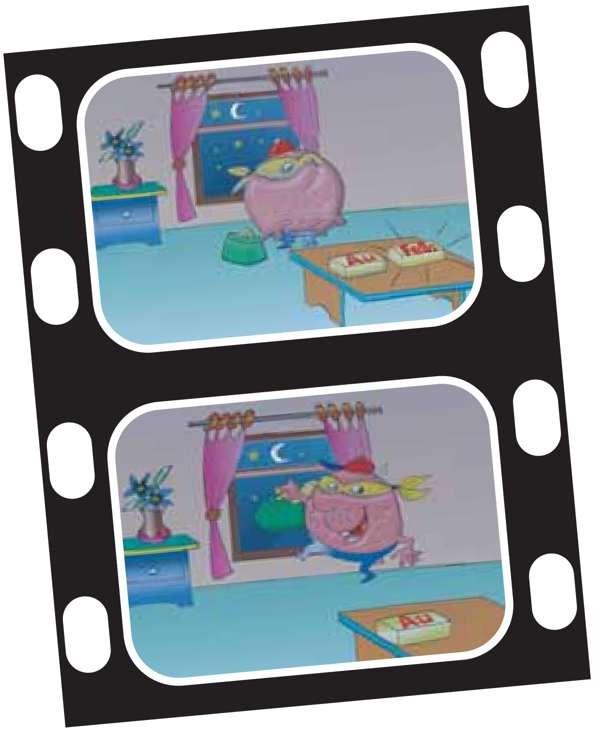


Group 8B

In the periodic table, elements in each column show similiar chemical and physical properties. However, in group 8B, similarities between the elements are observed within the same period. For example Fe, Co and Ni have these kinds of similarities.

These elements exhibit ferromagnetism and are strongly attracted by a magnetic field. They give two electrons from the outer orbital (4s), so they have +2 oxidation number in their compounds. Sometimes they can also give one of their electrons from the 3d orbital to have +3 oxidation number.

They are similar in having high densities and melting points, taking different oxidation numbers, forming complex compounds, having characteristic ionic colors and being not very active. In this group, the most active metal is Fe and the least active metal is nickel.



Preparation

a. In Industry

In metallurgy, iron is obtained in a blast furnace from ores of iron oxides by reduction of carbon. Molten iron and slag are withdrawn at the bottom of the furnace. The first step in the metallurgy of iron is usually roasting the ore to remove water, decompose carbonates and, especially, to oxidize sulfides.

$$Fe_2O_3 \cdot 2H_2O \xrightarrow{heat} Fe_2O_3 + 2H_2O\uparrow$$

$$FeCO_3 \xrightarrow{heat} FeO + CO_2\uparrow$$

$$2FeS + 3O_2 \xrightarrow{heat} 2FeO + 2SO_2$$

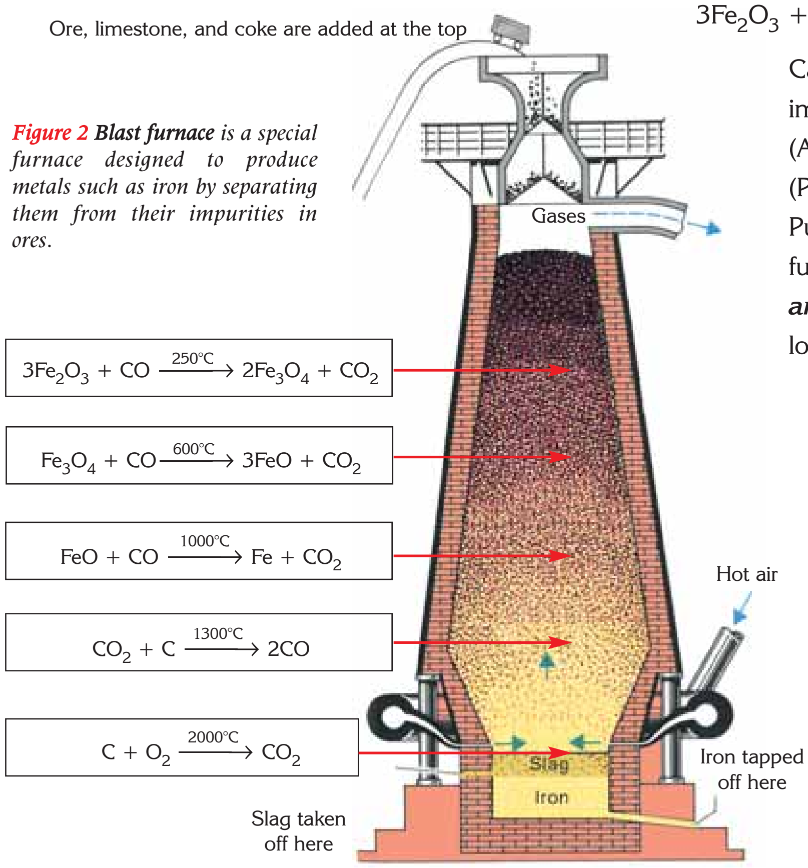


Figure 2 ***Blast furnace*** *is a special furnace designed to produce metals such as iron by separating them from their impurities in ores.*

The roasted ore, limestone flux and coke are charged continuously into the top of the furnace. At the same time, near the bottom of the furnace, preheated air enriched with oxygen is blown into the furnace. Therefore, coke and air produce carbon monoxide in two steps.

$$C(s) + O_2(g) \xrightarrow{2000°C} CO_2(g)$$

$$2C(s) + CO_2(g) \xrightarrow{1300°C} 2CO(g)$$

These are exothermic reactions so they increase the temperature of the furnace. The temperature at the bottom is about 2000°C and about 200°C at the top.

The CO produced reacts with iron oxide to reduce iron.

$$FeO + CO \longrightarrow Fe + CO_2$$

$$Fe_3O_4 + CO \longrightarrow 3FeO + CO_2$$

$$3Fe_2O_3 + CO \longrightarrow 2Fe_3O_4 + CO_2$$

Calcium carbonate, $CaCO_3$, added to remove impurities like sand (SiO_2), aluminum oxide (Al_2O_3), and tetraphosphorous pentoxide (P_4O_{10}), is decomposed into CaO and CO_2. Pure iron is removed at the bottom of the furnace. ***The iron obtained is called pig iron, and is not 100% pure***. It is brittle, but it has a low melting point.

A molten iron charge is added to a basic oxygen furnace.

b. In the Laboratory

Iron is prepared by the following methods in the laboratory:

1. H_2 gas is passed over hot iron (III) oxide.

$$Fe_2O_3(s) + 3H_2(g) \xrightarrow{\text{heat}} 2Fe(s) + 3H_2O(l)$$

2. Molten iron oxides are reduced by more active metals.

$$3FeO(l) + 2Al(s) \longrightarrow 3Fe(s) + Al_2O_3(s)$$

$$Fe_2O_3(l) + 2Al(s) \longrightarrow 2Fe(s) + Al_2O_3(s)$$

3. By the electrolysis of solutions of some iron salts.

$$FeCl_2(aq) \xrightarrow{\text{electrolysis}} Fe(s) + Cl_2(g)$$

The crucible contains a mixture of Fe_2O_3 and aluminum powder. A small amount of a mixture of potassium chlorate and sugar is placed on top of the Fe_2O_3 - Al mixture and a few drops of concentrated sulfuric acid are added to start the reaction.

The heat of the strongly exothermic reaction of H_2SO_4 with the sugar – $KClO_3$ mixture ignites the Fe_2O_3 – Al mixture which reacts violently in a strongly exothermic reaction. A shower of white hot sparks is emitted and the crucible becomes red hot.

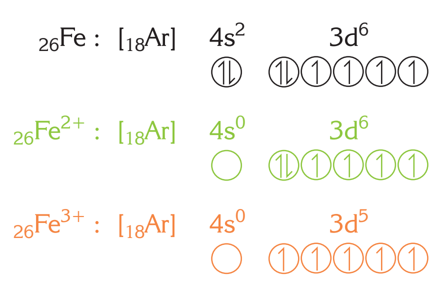
A ball of white hot iron can be seen glowing in the bottom of the crucible.

When the crucible has cooled, a magnet can be used to pick up the ball of iron.

Figure 3 *The Thermite Process: The reduction of iron(III) oxide to iron using aluminum as the reducing agent*

1.2 CHEMICAL PROPERTIES

Iron is a rather inactive metal with $[_{18}Ar]4s^23d^6$ electron configuration. In compounds, iron takes +2 and +3 oxidation states by giving 2 electrons from 4s and sometimes one more electron from 3d.



Reactions

1. Iron reacts with dilute solutions of strong acids.

$$Fe(s) + 2HCl(dil.) \longrightarrow FeCl_2(aq) + H_2(g)$$

$$Fe(s) + H_2SO_4(dil.) \longrightarrow FeSO_4(aq) + H_2(g)$$

The reactions of iron with oxidizing acids form its salts, containing Fe^{3+} ions

$2Fe(s) + 6H_2SO_4(conc) \longrightarrow Fe_2(SO_4)_3(aq) + 3SO_2 + 6H_2O$

$Fe(s) + 4HNO_3(dil.) \longrightarrow Fe(NO_3)_3(aq) + NO(g) + 2H_2O(l)$

Concentrated nitric acid causes a layer of Fe_3O_4 to form on the surface of the iron. This layer deactivates the metal.

$3Fe(s) + 8HNO_3(conc.) \longrightarrow Fe_3O_4(s) + 8NO_2(g) + 4H_2O(l)$

2. Iron produces mixed oxide of iron, Fe_3O_4 ($FeO \cdot Fe_2O_3$) by reacting with water vapor.

   $3Fe(s) + 4H_2O(g) \longrightarrow Fe_3O_4(s) + 4H_2(g)$

3. When iron is heated with sulfur iron sulfide, FeS forms

   $Fe(s) + S(s) \xrightarrow{t} FeS(s)$

4. At high temperature, it reacts with halogens.

   $2Fe(s) + 3Cl_2(g) \xrightarrow{1200°C} 2FeCl_3(s)$

5. Moisture and oxygen cause the formation of crystal hydrate of iron (III) oxide, $Fe_2O_3 \cdot nH_2O$.

   $4Fe(s) + 3O_2(g) + 2nH_2O(l) \longrightarrow 2Fe_2O_3 \cdot nH_2O(s)$

Iron metal reacts with HCl to release hydrogen gas.

Iron burns in oxygen

EXTRA

Corrosion

Iron and steel, the most commonly used materials corrode in many areas including most outdoor atmospheres. Usually they are selected not for their corrosion resistance, but for such properties as strength, ease of fabrication, and cost. These differences indicate the rate of metal lost due to rusting.

$4Fe(s) + 3O_2(g) + 2H_2O(l) \rightarrow \underset{\text{red brown}}{2Fe_2O_3 \cdot 2H_2O(s)}$

1.3 COMPOUNDS

Iron has +2 and +3 oxidation states in its compounds. Fe^{2+} ion is called ferrous and compounds that contain Fe^{2+} ion are called ferrous compounds. These compounds are obtained by the reaction of HCl and H_2SO_4 with metallic iron. Fe^{2+}, with green color, is easily oxidized to Fe^{3+}, which is colorless, by combining with oxygen in air. Fe^{3+} ion is called ferric and Fe^{3+} compounds are called ferric compounds. Such compounds can be obtained from the reaction of iron with fluorine, chlorine or dilute nitric acid.

a. Iron (II) compounds (Ferro Compounds)

1. Iron (II) chloride, $FeCl_2$

It is obtained by passing hydrogen chloride gas over heated iron. $FeCl_2$ is a white colored crystal.

$Fe(s) + 2HCl(g) \longrightarrow FeCl_2(s) + H_2(g)$

Iron (II) chloride in basic solutions produces iron (II) hydroxide solid that seems gelatinous.

$FeCl_2(aq) + 2NaOH(aq) \rightarrow Fe(OH)_2(s) + 2NaCl(aq)$ (figure 4).

2. Iron (II) sulfate, $FeSO_4 \cdot 7H_2O$

It is obtained by the reaction of dilute sulfuric acid with iron. Unhydrated iron (II) sulfate is colorless but when it is hydrated it takes green color. In air it is unstable.

$Fe(s) + H_2SO_4(dil) \longrightarrow FeSO_4(aq) + H_2(g)$

When iron (II) sulfate is heated it is decomposed into iron (II) oxides and sulfur oxides.

$2FeSO_4(s) \xrightarrow{heat} Fe_2O_3(s) + SO_2(g) + SO_3(g)$

3. Iron (II) oxide, FeO

This compound is produced by decomposition of iron (II) oxalate.

$$FeC_2O_4(s) \xrightarrow{heat} FeO(s) + CO(g) + CO_2(g)$$

FeO is also unstable in air.

$$4FeO(s) + O_2(g) \longrightarrow 2Fe_2O_3(s)$$

b. Iron(III) Compounds (Ferric Compounds)

The compounds of iron with +3 oxidation state are prepared by reactions of iron and chlorine gas directly, and of iron with concentrated H_2SO_4 or HNO_3.

1. Iron(III) chloride, $FeCl_3$

When iron is reacted with chlorine gas, it produces iron(III) chloride.

$$2Fe(s) + 3Cl_2(g) \xrightarrow{heat} \underset{\text{reddish-brown}}{2FeCl_3(s)}$$

If this compound reacts with water it produces a golden yellow colored crystal hydrate, $FeCl_3 . nH_2O$ compound.

$$FeCl_3(s) + nH_2O(l) \longrightarrow FeCl_3 . nH_2O(s)$$

2. Iron (III) hydroxide, $Fe(OH)_3$

It is obtained by the reaction of Fe^{3+} with a base or carbonates. It is similar to gelatin. $Fe(OH)_3$ is a reddish-brown colored precipitate which shows amphoteric property.

$$Fe^{3+}(aq) + 3KOH(aq) \rightarrow Fe(OH)_3(s) + 3K^+(aq)$$

When $Fe(OH)_3$ is heated, it is decomposed into iron (III) oxide and water.

$$2Fe(OH)_3 \xrightarrow{heat} Fe_2O_3 + 3H_2O$$

3. Iron (III) oxide, Fe_2O_3

In nature Fe_2O_3 is found in hematite and limonite minerals. It can be obtained by several methods.

$$2FeCl_3 + 3H_2O \xrightarrow{heat} Fe_2O_3 + 6HCl$$

$$4FeO + O_2 \longrightarrow 2Fe_2O_3$$

$$2Fe(OH)_3 \xrightarrow{heat} Fe_2O_3 + 3H_2O$$

$$4Fe(OH)_2 + O_2 \longrightarrow 2Fe_2O_3 + 4H_2O$$

The most common preparation method of Fe_2O_3 is the burning of pyrite, FeS_2 mineral.

$$4FeS_2 + 11O_2 \rightarrow 2Fe_2O_3 + 8SO_2$$

c. Iron(II, III) oxide, Fe_3O_4

Fe_3O_4, mixed oxide, is obtained by passing heated steam over iron metal or heating Fe_2O_3

$$3Fe + 4H_2O \xrightarrow{heat} Fe_3O_4 + 4H_2$$

$$6Fe_2O_3 \xrightarrow{heat} 4Fe_3O_4 + O_2$$

Fe_3O_4 is found in nature as black colored magnetite.

Figure 4 *Reactions of $Fe^{2+}(aq)$*

Left: When NaOH(aq) is added to a solution of a soluble Fe^{2+} salt, a dirty white precipitate of $Fe(OH)_2$ is formed. Center: In a few minutes the green color of $Fe(OH)_2$ starts to become a dark brown as the $Fe(OH)_2$ is oxidized by the air to $Fe(OH)_3$. Right: If aqueous hydrogen peroxide is added to the $Fe(OH)_2$ precipitate, it is immediately oxidized to red-brown $Fe(OH)_3$.

When an Fe^{2+} solution is added to a purple solution of potassium permanganate, it is immediately decolorized, as purple MnO_4^- is reduced to colorless Mn^{2+}. An aqueous solution of Mn^{2+} is actually pale pink, but here the solution is too dilute for the color to be seen.

Magnetite and compass

USES

Several transition metals are used in the fabrication of durable items. Iron is used in many different areas because of its cost, hardness and abundance in nature. Iron is the raw material from which steel is produced. Steel, the alloy of iron, is the basic material which is widely used in industry. Because it is strong, abundant and cheap, it is used in bridges, buildings, automobiles and airplanes.

Iron is used in the medical treatment of anemia. Since lack of iron in blood cells prevents the formation of hemoglobin, red blood cells do not function properly.

Rust Prevention

Method	Where it is used	Comment
A coat of paint	Ships,bridges, cars, other large objects (a)	If the paint is scratched, the exposed iron starts to rust. Corrosion can spread to the iron underneath intact paint.
A film of oil or grease	Moving parts of machinery, e.g. car engines	The film of oil or grease must be renewed frequently.
A coat of plastic	Kitchenware, e. g. draining rack	If the plastic is torn, the iron starts to rust.
Chromium plating	Kettles, cycle handle bars (b)	The layer of chromium protects the iron beneath it and also gives a decorative finish. It is applied by electroplating.
Galvanizing (zinc plating)	Galvanized steel girders are used in the construction of buildings and bridges (c)	Zinc is above iron in the reactivity series. Zinc will corrode in preference to iron. Even if the layer of zinc is scratched, as long as some zinc remains, the iron underneath does not rust. Zinc cannot be used for food cans because zinc and its compounds are poisonous.
Tin plating	Food cans	Tin is below iron in the reactivity series. If the layer of tin is scratched, the iron beneath it starts to rust.
Staninless steel	Cutlery, car accessories, e.g. radiator grille	Steel containing chromium (10-25%) does not rust.
Sacrificial protection	Ships (d)	Blocks of zinc are attached to the hulls of ships below the waterline. Being above iron in the reactivity series, zinc corrodes in preference to iron. The zinc blocks are sacrificed to protect the iron. As long as there is some zinc left, it protects the hull from rusting. The zinc blocks must be replaced after a certain time.

Table 2 *Rust prevention methods depending on application areas.*

(a) Paint protects the car from corrosion

(b) Chromium plating protects bicycle handlebars

(c) Galvanized steel girders

(d) Zinc bars protect the ship's hull

reading

The Manufacture of Steel

The term steel is a common name for many different alloys of iron. Steel is made from iron by removing impurities and adding substances such as manganese, chromium, nickel, tungsten, molybdenum, and vanadium to produce alloys with properties that make the material suitable for specific uses. Most steels also contain small but definite percentages of carbon (0.04 - 2.5%). Thus a large part of the carbon contained in iron must be removed in the manufacture of steel.

Steels can be classified in three main categories: (1) carbon steels, which are primarily iron and carbon; (2) stainless steels, low-carbon steels containing about 12% chromium; and (3) alloy steels, specialty steels that contain large amounts of other elements to impart special properties for specific uses. Of the 141 million tons of steel produced in the United States in a recent year, 124.6 million tons (88.4%) were carbon steels, 14.8 million tons (10.5%) were stainless steels, and 1.6 million tons (1.1%) were alloy steels.

The principal process used in the production of steel is the basic oxygen process, which utilizes a cylindrical furnace with a basic lining, such as magnesium oxides or calcium oxides. A typical charge is 80 tons of scrap iron, 200 tons of molten iron, and 18 tons of limestone (to form slag). A jet of high-purity (99.5%) oxygen is directed into the white-hot molten charge through a water-cooled lance. The oxygen produces a vigorous reaction that oxidizes the impurities in the charge. In the central reaction zone, temperatures reach a level close to the boiling point of iron. The entire steelmaking cycle is completed in one hour or less. Electrostatic precipitators clean the gases resulting from the furnace reactions, making the furnaces virtually smokeless. The steel produced is of extremely high and uniform quality.

Carbon is the most important alloying element in steel. It may be present in combination with iron as cementite, Fe_3C, or as crystals of graphite. The reaction of iron with carbon is reversible, and cementite is stable only at high temperatures. If cementite is cooled slowly, it decomposes to iron and graphite. Thus steel containing cementite in solid solution in iron must be made by quenching (quickly cooling) the hot metal in water or oil. The rate of decomposition of cementite is very slow at lower temperatures, so quick cooling does not provide enough time for the decomposition to occur. This steel is hard, brittle, and light-colored. If the metal is cooled slowly, the cementite decomposes and the carbon is deposited largely as separate crystals of graphite. The product is softer and more flexible and has a much higher tensile strength than does steel that has been cooled rapidly.

Certain materials called scavengers are added to iron in the manufacture of steel to remove impurities, especially oxygen and nitrogen, and thus improve the quality of the product. The most important scavengers are aluminum, ferrosilicon, ferromanganese, and ferrotitanium. They react with dissolved oxygen and nitrogen, forming oxides and nitrides, respectively, which are removed with the slag.

For every ton of steel produced, about 25-30 pounds of nonferrous metals are added or used as coatings. By the aproppriate choice of the number and percentages of these elements, alloy steel of widely varying properties can be manufactured. Some of the important alloy steels and their features are given below.

Name	Composition	Characteristic properties	Uses
Manganese steel	10 18% Mn	Hard, tough, resistant to wear	Railroad rails, safes, armor plate, rock-crushing machinery
Silicon steel	1/5% Si	Hard, strong, highly magnetic	Magnets
Duriron	12-15% Si	Resistant to corrosion, acids	Pipes, kettles, condensers, etc.
Invar	36% Ni	Low coefficient of expansion	Meter scales, measuring tapes, pendulum rods
Chrome-vanadium	1-10% Cr, 0.15% V	Strong, resistant to strains	Axles
Stainless steel	14-18% Cr, 7-9% Ni	Resistant to corrosion	Cutlery, instruments
Permalloy	78% Ni	High magnetic susceptibility	Ocean cables
High-speed steel	14-20% W or 6-12% Mo	Retain temper at high temperatures	High-speed cutting tools
Nickel steel	2-4% Ni	Hard and elastic, resistant to corrosion	Drive shafts, gears, cables

Table 3 *Types of steel.*

2. NICKEL

Nickel was isolated by Swedish chemist Baron A. F. Cronstedt from niccolite ore. Nickel takes its name from koppar nickel, meaning copper gin in Swedish. It shows the properties of all metals. It is the hardest metal of the Fe - Co - Ni triad. It is a magnetic element, like iron and cobalt. Its melting point is 1455 °C and boiling point is 2193°C. The density of nickel is 8.9 g/cm^3. It is a silvery white colored and very bright metal. When powdered it can absorb very large amounts of hydrogen, so it is used in many organic reactions as a catalyst to absorb hydrogen gas.

2.1 OCCURRENCE

Nickel makes up 0.008% of the earth's crust. It makes up 10% of molten magma. In nature it is found in nikelite (NiAs), millerite (NiS) and nickel glance (NiAsS) ores.

An Iron-nickel meteorite

A nickel mineral

Preparation

Preparation of pure nickel from its ores depends on the type of ore. Generally the amount of nickel in the ore is enriched by the floating method. This enriched ore is roasted to supply oxides of nickel.

$$2NiS + 3O_2 \longrightarrow 2NiO + 2SO_2$$

When these oxides are reduced with water gas (H_2 + CO) or with carbon, impure nickel is obtained.

$$NiO + H_2 \longrightarrow Ni + H_2O$$

$$NiO + CO \longrightarrow Ni + CO_2$$

$$NiO + C \longrightarrow Ni + CO$$

Nickel is then purified. For this purpose, electrolysis or the Mond method are used. In the Mond method, impure nickel is treated with carbon monoxide to give nickel carbonyl, $Ni(CO)_4$ complex, at 50°C. Pure nickel can be obtained by decomposition of this compound at 200°C.

$$Ni(CO)_4 \xrightarrow{heat} Ni(s) + 4CO(g)$$

2.2 CHEMICAL PROPERTIES

Nickel with $[Ar]3d^8 4s^2$ electron configuration has +2 and +3 oxidation states. It is inactive in chemical reactions because of the oxide layer formed on the metal. It is very stable in humidity and water.

$_{28}Ni$: [Ar] $4s^2$ $3d^8$
⇅ ⇅⇅⇅↑↑

Reactions

1. Acids act on nickel slowly. It can be dissolved in dilute oxidizing acids quickly. However, concentrated HNO_3 deactivates on Ni metal because of the nickel oxides on the surface of the metal.

$$3Ni + 8HNO_3(dil.) \rightarrow 3Ni(NO_3)_2 + 2NO\uparrow + 4H_2O$$

$$Ni + HNO_3(conc.) \rightarrow \text{no reaction}$$

2. Nickel reacts with most of the nonmetals at high temperatures.

$$2Ni + O_2 \xrightarrow{t} 2NiO$$

$$Ni + Cl_2 \xrightarrow{t} NiCl_2$$

$$2Ni + C \xrightarrow{t} Ni_2C$$

3. At normal conditions nickel reacts with different acids.

$$Ni + H_2SO_4 \longrightarrow NiSO_4 + H_2\uparrow$$

$$Ni + 2HCl \longrightarrow NiCl_2 + H_2\uparrow$$

4. Nickel ions can form complexes with water, ammonia, cyanide, carbon monoxide and also with many different organic compounds.

$Ni^{2+} + 6H_2O \longrightarrow Ni(H_2O)_6^{2+}$

$Ni^{2+} + 6NH_3 \longrightarrow Ni(NH_3)_6^{2+}$

2.3 COMPOUNDS

Generally Ni has +2 oxidation state in its compounds, but sometimes it also takes +1, +3 and +4 oxidation states.

Nickel (II) ions in aqueous solutions are green, whereas unhydrated nickel (II) salts are yellow.

1. Nickel (II) hydroxide, $Ni(OH)_2$

It is a light green solid obtained by addition of bases into solutions of nickel (II) salts.

$Ni^{+2}(aq) + 2OH^-(aq) \longrightarrow Ni(OH)_2(s)$

Nickel (II) oxide reacts with strong oxidizing agents like ClO^-, Cl_2 or H_2O_2 in basic media. The product of these reactions is dark colored nickel (IV) oxide, NiO_2.

$Ni(OH)_2(s) + ClO^-(aq) \rightarrow NiO_2(s) + Cl^-(aq) + H_2O(s)$

Nickel (II) hydroxide dissolves in ammonia solution by forming a complex compound.

$Ni(OH)_2 + 6NH_3 \longrightarrow [Ni(NH_3)_6](OH)_2$

2. Nickel (II) oxide, NiO

Nickel (II) oxide is obtained by thermal decomposition of nickel carbonate or nickel (II) hydroxide.

$Ni(OH)_2 \xrightarrow{heat} NiO + H_2O$

NiO is a dark green, jelly-like, solid substance used in the ceramic and glass industries as green dye pigment.

3. Nickel (II) sulfate, $NiSO_4 \cdot 7H_2O$

One of the most important salts of nickel is nickel (II) sulfate, a green colored solid. It is obtained by the reaction of $NiCO_3$ with dilute sulfuric acid.

$NiCO_3(s) + H_2SO_4(aq) \rightarrow NiSO_4(s) + CO_2(g) + H_2O(l)$

Nickel (II) sulfate solution is used as electrolyte in plating other metals with nickel. This process is called nickel plating. It is stronger than other plating materials.

USES

Nickel is used in industry in different alloys. Some important alloys and their properties are as follows.

1. An alloy containing 67% nickel and 28% copper is used to prevent corrosion in propellers, etc. It is also used in space technology.
2. Nicrome, which contains 60% Ni, 15% Cr and 25% Fe, and chromel, which contains 90% Ni, and 10% Cr, are used in electrical wires with high resistance.

   Ni-Cr alloys have a high resistance to extreme cold and heat so they are used in steel manufacture for knives, spoons and for other kitchen tools, and for laboratory instruments.
3. Al - Ni - Co alloy is used in telephone wires, and in wireless and radio receivers because of its magnetic property.
4. Ni - Cu - Zn alloy is used in brilliants, jewelry and coins.

An apparatus coated with Ni.

3. CHROMIUM

Chromium is the first member of group 6B. Chromium was discovered by the French scientist L. N. Vauquelin in 1797. The name chromium comes from the Latin word chroma, meaning color.

Chromium compounds are used in the dye industry as a pigment. Chromium ions give the bright and attractive colors to precious stones such as ruby and sapphire. Pure chromium is grey in color, hard and bright like silver. The melting point is 1907°C, the boiling point is 2671°C and its density is 7.19 g/cm^3 at room temperature.

3.1 OCCURRENCE

The percentage of chromium is about 0.14% by mass in the earth's crust. It generally exists in volcanic stones and meteors. The most important mineral of chromium is chromite ($FeO . Cr_2O_3$), which has a brownish-black color. Moreover, its rarely found ore is chrocoite ($PbCrO_4$).

Chromite

Preparation

In industry, elemental chromium is obtained from the reduction of chromite mineral by coal or aluminum in an electric furnace.

$$FeO \cdot Cr_2O_3 + 4C \longrightarrow 2Cr + Fe + 4CO \quad \text{or}$$

$$3(FeO \cdot Cr_2O_3) + 8Al \longrightarrow 6Cr + 3Fe + 4Al_2O_3$$

At the end of the reaction a small amount of chromium carbide is produced.

Pure chromium is obtained from the reduction of Cr_2O_3 by Al or Si.

$$Cr_2O_3 + 2Al \longrightarrow 2Cr + Al_2O_3$$

$$2Cr_2O_3 + 3Si + 3CaO \longrightarrow 4Cr + 3CaSiO_3$$

Very pure chromium is prepared by the electrolysis of solutions of chromium salts.

3.2 CHEMICAL PROPERTIES

Chromium is in group 6B in the periodic table, and has electron configuration $[_{18}Ar]4s^13d^5$.

$_{24}Cr$:	[Ar]	$4s^1$	$3d^5$
		①	①①①①①

That configuration is more stable than $[_{18}Ar]4s^23d^4$ due to its spherical geometry (half-filled orbitals) structure. The main oxidation states of chromium are +2, +3 and +6, but it may exist from +1 to +6 oxidation states.

Powdered chromium is more active. It may be reacted easily with NO_3^- and SO_4^{2-} compounds, and with O_2 gas. Its surface is then coated with a thin oxidized layer which prevents further reaction. That is why chromium is not affected by air and water.

Reactions

1. Chromium metal reacts with halo-acids, such as HCl and HBr, slowly. But if we heat the medium, it reacts quickly, depending on the concentration of acids, to produce H_2 gas. Light blue colored Cr^{2+} salt solutions are produced.

$$Cr + 2HCl(dil.) \longrightarrow \underset{\text{pale blue}}{CrCl_2} + H_2$$

2. Hot and concentrated sulfuric acid reacts with chromium to form SO_2 gas.

$$2Cr + 6H_2SO_4(conc.) \rightarrow Cr_2(SO_4)_3 + 3SO_2 + 6H_2O$$

Chromium reacts with dilute H_2SO_4 to give $CrSO_4$ salt and H_2 gas.

$$Cr + H_2SO_4(dil.) \longrightarrow CrSO_4 + H_2$$

3. Chromium does not react with dilute and concentrated HNO_3 at room temperature because of the formation of a protective oxide layer on the surface of the metal. Chromium metal becomes passive in the acid solutions mentioned.

$$Cr + HNO_3 \longrightarrow \text{no reaction}$$

4. Heated chromium metal reacts with some nonmetals to produce the following Cr(III) compounds:

$$2Cr + 3Cl_2 \xrightarrow{1200°C} 2CrCl_3 \text{ (Black)}$$

$$4Cr + 3O_2 \xrightarrow{600°C} 2Cr_2O_3 \text{ (Green)}$$

$$2Cr + N_2 \xrightarrow{900°C} 2CrN \text{ (Black)}$$

3.3 COMPOUNDS

Cr^{3+} and Cr^{2+} compounds are stable in solutions and Cr^{2+} compounds are thus fairly reducing agents. Cr^{2+} ions give blue color and Cr^{3+} ions give green color.

The most important compounds of chromium with oxygen are chromium (III) oxide (Cr_2O_3), chromium (VI) oxide (CrO_3), and the chromate (CrO_4^{2-}) and dichromate ($Cr_2O_7^{2-}$) salts.

1. Chromium (III) Oxide, Cr_2O_3

Chromium (III) oxide is a green colored powder.

It is used as a pigment. It is a carcinogen!

Cr_2O_3 is obtained by heating chromium (III) hydroxide or ammonium dichromate.

$$2Cr(OH)_3 \xrightarrow{heat} Cr_2O_3 + 3H_2O \quad \text{or}$$

$$(NH_4)_2Cr_2O_7 \xrightarrow{heat} Cr_2O_3 + N_2 + 4H_2O$$

(Figure 5)

Cr_2O_3 is an amphoteric oxide. That is why it reacts with both acids and bases to give salt and H_2O.

$$Cr_2O_3 + 6HCl \longrightarrow 2CrCl_3 + 3H_2O$$

$$Cr_2O_3 + 2NaOH \xrightarrow{heat} 2NaCrO_2 + H_2O$$

Figure 5 Decomposition of ammonium dichromate, volcano reaction. (This reaction must be carried out carefully! Because product is a very dangerous oxide.

2. Chromium (VI) Oxide, CrO_3

Chromium (VI) oxide is a red colored solid that melts at 197°C. It is a powerful oxidizing agent. Therefore it is used to clean the glassware in laboratories.

When chromate or dichromate salts are reacted with concentrated H_2SO_4 solution, CrO_3 forms.

$$K_2Cr_2O_7(s) + H_2SO_4(conc.) \longrightarrow 2CrO_3(s) + K_2SO_4(aq) + H_2O(l)$$

$$K_2CrO_4(s) + H_2SO_4(conc.) \longrightarrow CrO_3(s) + K_2SO_4(aq) + H_2O(l)$$

If chromium(VI) oxide is heated above 197°C (its melting point), it decomposes to chromium (III) oxide and oxygen gas.

$$4CrO_3(s) \xrightarrow{\text{heat}} 2Cr_2O_3(s) + 3O_2(g)$$

CrO_3 reacts with H_2O to give $H_2Cr_2O_7$ and H_2CrO_4 acids.

$$2CrO_3 + H_2O \longrightarrow H_2Cr_2O_7$$

3. Chromates (CrO_4^{2-}) and dichromates ($Cr_2O_7^{2-}$)

Chromates of alkali metals, magnesium and calcium are soluble in water. Soluble chromates have usually yellow color.

Soluble chromates are used to detect some ions in analytical chemistry.

$$Na_2CrO_4 + Pb(NO_3)_2 \longrightarrow PbCrO_4\downarrow + 2NaNO_3$$

Chromates react with acids and turn to orange colored dichromates.

$$2CrO_4^{2-} + 2H^+ \rightleftharpoons Cr_2O_7^{2-} + H_2O$$

If base is added to the given equilibrium, the reaction reverses. In this way, CrO_4^{2-} ions are obtained again.

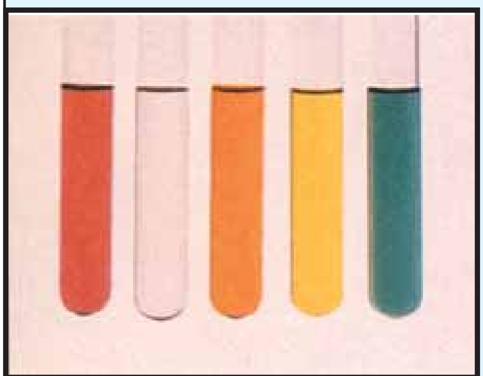
Chromium compounds are used to detect alcohol level in blood.

USES

Chromium containing alloys are used in coins.

In general, chromium metal is not used alone. Most chromium is used in the manufacture of stainless steel. It forms important alloys with iron, nickel, manganese and aluminum, such as ferro-chromium (60-70% Cr and 30-40% iron) and nichrome (60%Ni, 25%Fe and 15%Cr), which has very high resistance to heat, and is used to build electrical furnaces.

Chromium is used for coating metals against corrosion. Chromium powder is used in the process of tanning, in dyes and in textiles.



Solutions of chromium compounds have a variety of colors.

4. TITANIUM

Titanium is the first member of group 4B. Titanium was discovered by William Gregor in 1791.

The meaning of titanium is "son of the earth" in Greek mythology. If we compare the physical properties of metals, the boiling points of the metals in group 4B are higher than other metals, but the melting points are lower. Titanium is a silvery metallic colored solid at room temperature. The melting point of titanium is 1668°C and the boiling point is 3287°C. The density of titanium is 4.5 g/cm^3.

Titanium is very hard and has high resistance to corrosion. It is ductile and a good conductor of heat and electricity.

4.1 OCCURRENCE

The amount of titanium in the earth's crust is 0.58%. The main minerals of titanium element are ilmenite ($FeTiO_3$) and rutile (TiO_2). Because of its tendency to react with carbon, oxygen and nitrogen, it is very difficult to obtain pure titanium.

Preparation

Generally, pure titanium is prepared by the following steps.

a. First TiO_2 is reacted with carbon and chlorine gas at about 900°C, to obtain $TiCl_4$.

$$2TiO_2 + 3C + 4Cl_2 \xrightarrow{heat} 2TiCl_4 + 2CO\uparrow + CO_2\uparrow$$

b. After that, $TiCl_4$ is evaporated at about 800°C in the absence of oxygen. Then it reacts with magnesium to give titanium metal.

$$TiCl_4 + 2Mg \longrightarrow Ti + 2MgCl_2 + 481\ kJ.$$

c. Finally, titanium metal is purified by reaction with nitric acid (2%).

4.2 CHEMICAL PROPERTIES

The electron configuration of titanium is [Ar] $3d^2 4s^2$. Titanium may exist in +2, +3 or +4 oxidation states. Ti^{3+} and Ti^{4+} are more stable, and Ti^{2+} and Ti^{3+} compounds are colorful and paramagnetic. The most important compounds are TiO_2, TiCl, TiC and TiN.

$_{22}Ti$: [Ar]	$4s^2$	$3d^2$
	⇅	↑ ↑ ◯ ◯ ◯

Reactions

The elements in group 4B are generally good reducing agents and do not react with concentrated acids at room temperature. Titanium reacts with all nonmetals under proper conditions. Some reactions of titanium are performed at high temperature.

1. Titanium reacts with halo-acids to form halides of titanium.

$$2Ti + 6HF \longrightarrow 2TiF_3 + 3H_2\uparrow$$

$$2Ti + 6HCl \longrightarrow 2TiCl_3 + 3H_2\uparrow$$

2. At high temperature, titanium reacts with sulfuric acid.

$$2Ti + 6H_2SO_4(conc.) \xrightarrow{heat} Ti_2(SO_4)_3 + 3SO_2 + 6H_2O$$

3. It reacts with hot water by releasing hydrogen gas.

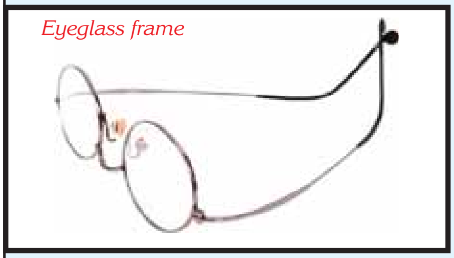
$$Ti + 2H_2O \longrightarrow TiO_2 + 2H_2\uparrow$$

EXTRA

Why is titanium dioxide the most widely used white pigment?

Titanium dioxide has become the predominant white pigment in the world because of its high refractive index, lack of absorption of visible light, ability to be produced in the right size range, and its stability. It is the whitest known pigment, unrivalled for color, opacity, stain resistance, and durability. It is also non-toxic. The main consuming industries are paint, printing inks, plastics, and ceramics.

USES



Eyeglass frame

Titanium is alloyed with aluminum, molybdenum, manganese, iron and other metals. These alloys are used in airframes and engines. It has excellent resistance to sea water and is used to prevent corrosion of ships.

Some food and water tanks are made from titanium alloys. Titanium alloys are also used in the manufacture of eyeglass frames, pens, cameras, tennis rackets, golf clubs and bicycles.

Because of its resistance to corrosion, it is also used in the production of medical equipment and prostheses.

Titanium paint is an excellent reflector, and is used in solar observatories. TiO_2 is used as a white pigment.

Titanium compounds are used in the manufacture of paints.

Without titanium metal, modern planes and the space industry could not be imagined.

Tennis racket

5. ZINC

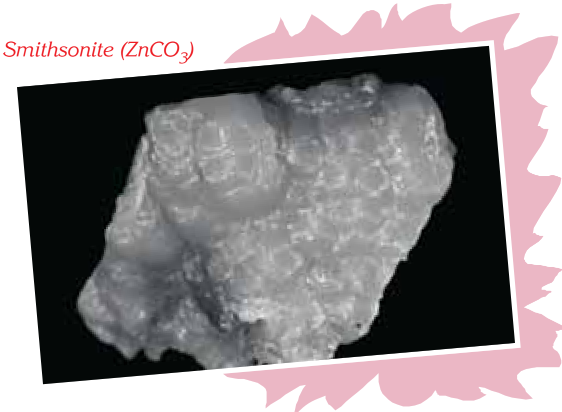
Zinc is the first member of group 2B. Its electron configuration is $[Ar]3d^{10}4s^2$. Zinc takes +2 oxidation state in its compounds.

Zinc is a bluish-white metal with melting point of 419.5°C and boiling point of 907°C. The density of zinc is 7.14 g/cm^3. At room temperature, it is fairly hard and brittle. But, if it is heated to 100-150°C, it becomes a soft metal that can easily be hammered to give shape. It again becomes brittle over 200°C.

5.1 OCCURRENCE

Zinc is a relatively active metal in comparison with others, and is not found in elemental form in nature. It is found as compounds, such as zincblende (ZnS), willemite ($Zn_2SiO_4 \cdot H_2O$), smithsonite or calamine ($ZnCO_3$), and franklinite ($ZnO \cdot Fe_2O_3$) in crustal rocks.

Smithsonite ($ZnCO_3$)



Preparation

Zinc metal is readily available commercially so it is not normally necessary to make it in the laboratory. Most zinc production is based upon sulfide ores like zincblende (ZnS). The ore is enriched by some methods (such as filtration) and then rich ore is roasted in the blast furnaces.

$$2ZnS(s) + 3O_2(g) \xrightarrow{\text{heat}} 2ZnO(s) + 2SO_2(g)$$

After that, the ZnO obtained is reduced with carbon to form metallic zinc.During these processes, the temperature of the furnace must be above the boiling point.

$$ZnO + C \longrightarrow Zn + CO \qquad \Delta H = -232 \text{ kJ}$$

Occurred CO also reacts with ZnO to produce Zn metal.

$$ZnO + CO \longrightarrow Zn + CO_2 \qquad \Delta H = -69 \text{ kJ}$$

Zinc obtained can be evaporated and collected in containers as a liquid, by distillation.

Another method in the preparation of zinc is the electrolytic method. In this method, first crude zinc oxide is treated with sulfuric acid and impure zinc sulfate solution is obtained.

$$ZnO + H_2SO_4 \longrightarrow ZnSO_4 + H_2O$$

Next, the solution is electrolyzed and pure zinc is collected at the cathode.

$$2ZnSO_4 + 2H_2O \xrightarrow{\text{electrolysis}} 2Zn\downarrow + O_2 + 2H_2SO_4$$

5.2 CHEMICAL PROPERTIES

Because of its electron configuration ([Ar]$3d^{10}4s^2$), zinc takes only +2 oxidation state by giving two 4s electrons.

$_{30}$Zn : [Ar]	$4s^2$	$3d^{10}$
	⥮	⥮⥮⥮⥮⥮

Zinc is an amphoteric metal so it reacts with acids and bases. In dry air, zinc doesn't tarnish. In moist air, the surface of the metal tarnishes gray, but it doesn't oxidize further.

Reactions

1. Zinc metal burns in air with a greenish-blue flame to form white zinc oxide.

$$2Zn(s) + O_2(g) \xrightarrow{\text{heat}} 2ZnO(s)$$

The reaction of zinc with oxygen

2. Zinc reacts with halogens to form white colored halides.

$$Zn(s) + Cl_2(g) \xrightarrow{\text{heat}} ZnCl_2(s)$$

$$Zn(s) + I_2(g) \xrightarrow{\text{heat}} ZnI_2(s)$$

3. Zinc reacts with sulfur when heated, but no reaction of zinc with nitrogen is observed.

$$Zn(s) + S(s) \xrightarrow{\text{heat}} ZnS(s)$$

$$Zn(s) + N_2(g) \longrightarrow \text{No reaction}$$

4. Zinc reacts with dilute acids, except HNO_3, to give hydrogen gas.

$$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

$$Zn(s) + H_2SO_4(dil.) \longrightarrow ZnSO_4(aq) + H_2(g)$$

When oxidizing acids react with zinc, different gases are evolved depending on the type and concentration of the acid.

$4Zn(s) + 5H_2SO_4(conc.) \longrightarrow 4ZnSO_4(aq) + H_2S\uparrow + 4H_2O$

$Zn(s) + 4HNO_3(conc.) \longrightarrow Zn(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$

$Zn(s) + HNO_3(dil.) \longrightarrow Zn(NO_3)_2(aq) + N_2O(g) + H_2O(l)$

5. Reaction of Zn with strong bases.

$Zn + 2NaOH(conc.) + 2H_2O \longrightarrow Na_2[Zn(OH)_4](aq) + H_2(g)$

USES

Zinc is used in alloys such as brass and bronze, which are valued for their resistance to corrosion. It is also used in the galvanizing, protective coating, of other metals. Galvanized products include roofs, fences and gutters.

Zinc oxide is used as a pigment (chinese white), as a mild antiseptic, and in the treatment of skin irritations, such as diaper rash. It is also used in the manufacture of floor coverings, printing inks, textiles, electrical equipment, storage batteries and soap.

Some zinc compounds are phosphorescent and so are used in the manufacture of television tubes and X- ray screens. Zinc sulfate is used to check bleeding, and zinc chloride, being deliquescent, is used as a dehydrating agent.

6. CADMIUM

Cadmium is also one of the transition elements like iron, nickel and zinc. It has an electron configuration of $[Kr]4d^{10}5s^2$ in group 2B.

$_{48}Cd$: [Kr] $5s^2$ $4d^{10}$

⇅ ⇅⇅⇅⇅⇅

Cadmium is softer than zinc and is malleable and ductile. Although the chemical tendencies of cadmium are similar to those of zinc, it is less active than zinc. The most important property of Cd is to form low melting point alloys with tin (Sn) and bismuth (Bi).

6.1 OCCURRENCE

0.15 ppm of the earth's crust is elemental cadmium. It is usually found as its carbonates ($CdCO_3$) and sulfides (CdS). It is produced as a by-product in the production of zinc. Zinc and cadmium are separated by fractional distillation of their oxides. Cadmium oxide is more easily oxidized than zinc oxide. Because of its low boiling point, pure cadmium is separated first from the molten mixture by distillation.

6.2 CHEMICAL PROPERTIES

Chemically, cadmium is like Zn in many respects. For example, it dissolves in nonoxidizing acids to form 2+ charged ions. That is the only oxidation state of Cd in compounds.

Reactions

Free cadmium metal is moderately active. It acts as a reducing agent. At room temperature it reacts with acids.

$Cd + 2HCl(dil.) \longrightarrow CdCl_2 + H_2\uparrow$

$Cd + 4HNO_3(conc.) \rightarrow Cd(NO_3)_2 + 2NO_2\uparrow + 2H_2O$

At higher temperatures it reacts with steam and some non-metals.

$$Cd + H_2O \xrightarrow{heat} CdO + H_2\uparrow$$

$$2Cd + O_2 \xrightarrow{heat} 2CdO$$

$$Cd + Cl_2 \xrightarrow{heat} CdCl_2$$

$$Cd + S \xrightarrow{heat} CdS$$

6.3 COMPOUNDS

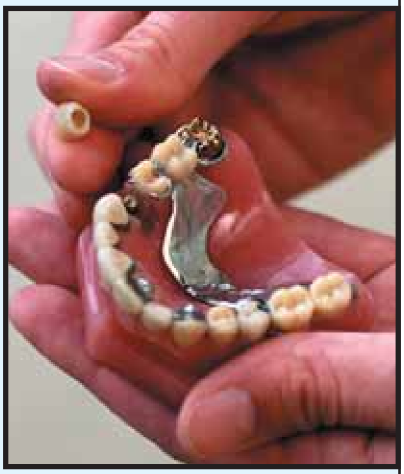
The most important compounds of cadmium are CdO, $Cd(OH)_2$, CdS, $CdSO_4$ and $CdCl_2$. Cadmium compounds are toxic. If absorbed by the body they can cause high blood pressure, heart disease and even painful death.

USES

Its chief use is as a protective coating on other metals and for making Ni-Cd batteries. Amalgams of cadmium are used in dental fillings. Cadmium oxide (CdO) is used in plating by electrolysis and as a catalyst in some Ni-Cd cell reactions.

Cadmium sulfide (CdS) is used in solar cells, photography, photocopiers and dyes. Cadmium sulfate ($CdSO_4$) is used in electroplating and in standard voltaic cells. Furthermore, cadmium is used in various areas such as the protection of iron against corrosion, and in aluminum solders. Because of its ability to capture electrons, it is used in nuclear reactors.

Cadmium poisoning is a painful bone disease. This illness was observed first in Japan in rice fields where field waters mixed with cadmium-containing zinc mineral. The illness was observed in people who ate the rice. This illness is harmful to the liver, causes kidney deficiencies, and lung illnesses. In addition, cadmium is thought to be one of the causes of hypertension.



Amalgams are the alloys of mercury with other metals. Cadmium is added to amalgams to stabilize their color.

SUPPLEMENTARY QUESTIONS

IRON

1. How can we explain the oxidation states of iron by using its electron configuration?

2. What is the meaning of ferromagnetic? Which elements in the periodic table show ferromagnetism?

3. Complete the following reactions, and balance them.

a. $Fe_3O_4(s) + H_2(g) \longrightarrow$

b. $Fe_2O_3(s) + Mg(s) \longrightarrow$

c. $FeCl_2(l) \xrightarrow{elect.}$

d. $Fe_2O_3(s) + C(s) \longrightarrow$

e. $Fe_2O_3(s) + CO(g) \longrightarrow$

f. $FeCl_2(aq) + Zn(s) \longrightarrow$

4. In metallurgy, what is the purpose of roasting the ore?

5. If an iron ore contains $FeCO_3$, Na_2CO_3, and FeS_2, what kind of reactions may form during roasting? Write their equations.

6. Complete and balance these reactions.

a. $Fe + HCl \longrightarrow$

b. $Fe_2O_3 + HNO_3 \longrightarrow$

c. $FeCl_2 + Cl_2 \longrightarrow$

d. $Fe(OH)_3 + KCl \longrightarrow$

e. $FeSO_4 + Ba(NO_3)_2 \longrightarrow$

f. $Fe(OH)_2 + H_2SO_4 \longrightarrow$

7. How can we differentiate $FeCl_2$ and $FeCl_3$ solutions? Suggest an easy way.

8. Which of the given compounds is/are thermally stable?

$FeSO_4$, $Fe(OH)_3$, $FeCO_3$, FeC_2O_4

9. What is corrosion? Research the prevention methods, used in the home, against corrosion.

10. Write 2 reactions in which Fe^{2+} ions oxidize to Fe^{3+} ions.

11. Give the colors of the following substances.

a. FeO	**b.** Fe_2O_3	**c.** Fe_3O_4
d. FeS	**e.** FeS_2	**f.** $Fe(OH)_2$
g. $Fe(OH)_3$	**h.** $FeCl_2$	**i.** $FeCl_3$
j. $FeSO_4$	**k.** $FeSO_4 . 7H_2O$	

12. Write a few uses of iron in industry.

13. What is anemia? How can it be treated?

14. Sodium hydroxide is added to a mixture of iron (II) chloride. Write the reaction equation.

15. A mineral contains 72.36% iron and 27.64% oxygen in its structure. Find the chemical formula of this iron oxide.

16. What is the difference in mass of iron plate immersed into a copper sulphate solution, if as a result of the reaction, 20.8 g copper metal is added to the iron plate?

17. How many grams of iron (II) sulphate are obtained as a result of reacting 14 g iron with diluted sulphuric acid?

18. How many grams of iron (III) sulphate are obtained as a result of reacting 120 g of 10% NaOH hydroxide solution with excess iron (III) chloride?

19. What are the main ores of iron?

20. Fill in the blanks.

a. ________________ is the most important element for steel.

b. The d-block elements are also called ________________.

c. The chemical formula of rust is ________________.

d. Ferromagnetic elements are ________________.

21. Write the production reaction of iron (II) chloride and iron (III) chloride.

22. Complete and balance these reactions.

a. $Fe + HCl \longrightarrow$

b. $Fe_2O_3 + HNO_3 \longrightarrow$

c. $Fe_2O_3 + H_2O \longrightarrow$

d. $FeCl_2 + Cl_2 \longrightarrow$

e. $FeSO_4 + Ba(NO_3)_2 \longrightarrow$

f. $Fe(OH)_2 + H_2SO_4 \longrightarrow$

NICKEL

23. Find the oxidation states of Ni atom in the given compounds.

a. NiO **b.** $NiSO_4 . 7H_2O$ **c.** $Ni(CO)_4$ **d.** $Ni(NH_3)_6^{2+}$

24. Write the correct reaction of nickel oxide with:

a. Carbon **b.** Water-gas

25. Which compounds of nickel are the most important? Why?

26. Complete the following reactions.

a. $Ni(OH)_2 + HCl \longrightarrow$

b. $Ni(OH)_2 + NH_3 \longrightarrow$

c. $Ni(OH)_2 + H_2O_2 \longrightarrow$

27. What are industrial uses of nickel?

ZINC

28. What is the electron configuration of Zn metal?

29. Describe the characteristic properties of zinc.

30. Write the main zinc ores.

31. Generally zinc is found in the form of compounds in the earth's crust and in volcanic rocks. Why is zinc not found in elemental form?

32. How is zinc extracted from ores?

33. Which react(s) with zinc metal to produce hydrogen gas?

a. Steam

b. Ammonia solution

c. Dilute hydrochloric acid

d. Concentrated sulphuric acid

e. Sodium hydroxide solution

34. Write the equations of the given chemical reactions.

a. Zinc is burned in air.

b. Zinc is heated with sulfur.

c. Zinc sulfide is heated in a stream of oxygen gas.

CADMIUM

35. Write the electron configuration of Cd.

36. What is the oxidation state of cadmium in compounds? Explain.

37. What is the main difference between cadmium and zinc?

38. Write the similarities of cadmium and zinc.

39. What is cadmium poisoning?

40. Where do we use cadmium and its compounds?

41. Research Ni-Cd batteries. Prepare a report on this topic.

42. Research why cadmium compounds are toxic. Decide whether used Ni-Cd batteries are dangerous for the environment or not. Explain, with reasons.

43. Where do we use cadmium?

44. **a.** Describe some characteristics and components of steel.

b. Discuss various steps and chemical reactions involved in the manufacture of steel.

45. Fill in the blanks.

a. Zinc, with an electron configuration of 2, 8, 18, 2 has an atomic number and a valency of ________________

46. Fill in the blanks.

At one stage in the industrial extraction of zinc, ZnO is reduced by heating with ________________ to give liquid zinc and ____________ gas.

CHROMIUM

47. Why was the name chromium given to the element with atomic number 24? Collect useful information.

48. Explain the industrial preparation of chromium.

49. Write the reactions of chromium with the following substances (if possible).

a. Dilute HBr **b.** Dilute HNO_3 **c.** Concentrated HNO_3

d. Dilute H_2SO_4 **e.** Cl_2 **f.** S_8

g. C **h.** P_4

50. Write the color of the following chromium compounds.

a. Cr_2O_3 **b.** CrO_3 **c.** Na_2CrO_4

d. $Na_2Cr_2O_7$ **e.** $(NH_4)_2Cr_2O_7$ **f.** $PbCrO_4$

51. Which property of chromium allows it to be used in alloys?

52. Find ten examples of chromium plated things in your daily life.

TITANIUM

53. Write the most important ores of titanium.

54. Why is titanium dioxide widely used as a white pigment?

55. Write the electron configurations of Ti^{2+} and Ti^{4+} ions and decide whether they are paramagnetic or diamagnetic.

56. Complete the given reactions.

a. $TiCl_4 + Mg \longrightarrow$ **b.** $TiO_2 + Al \longrightarrow$

c. $TiO_2 + C + Br_2 \longrightarrow$ **d.** $Ti + H_2SO_4 \longrightarrow$

57. Which property of titanium allows it to be used in alloys? Write a few applications of titanium alloys.

MULTIPLE CHOICE QUESTIONS

1. Which one of the following is an ore of zinc?
A) Zinc blend B) Albumine C) Smithsonite D) Zincit E) Ztince Kalamin

2. Which one does not contain zinc metal?
A) Zinc sulfate B) Brass C) Bronze D) Zinc oxide E) Hematite

3. Which one is not a property of zinc?
A) Zinc is an active metal.
B) Zinc Blend is the common ore of zinc.
C) Generally, zinc is found in pure metallic form in nature.
D) Charge of zinc is always 2+ in its compounds
E) CO is used to produce pure Zn metal from ZnO

4. Which of the following metals is in the same group with Fe?
A) Al B) Ba C) Na D) Ni E) Li

5. Which of the following is one of the oxidation numbers of Fe?
A) +1 B) +2 C) –2 D) +4 E) +5

6. Which of the following is not a natural form of Fe?
A) FeS_2 B) $FeCO_3$ C) Fe_2O_3 D) Fe_3O_4 E) $FeBr_2$

7. Lack of which element is the reason for anemia?
A) Fe B) Al C) Ca D) Sn E) Pb

8. Which element is the basic element for manufacturing steel?
A) Fe B) Br_2 C) Mg D) C E) Zn

9. Which one of the following is a product of the reaction:
$Ni(CO)_4 \xrightarrow{heat}$
A) CO_2 B) CO_3 C) CO D) NiO E) $NiCO_3$

10. Which of the following is the most common oxidation number of Ni?
A) +1 B) –1 C) +3 D) +4 E) +2

11. Which one of the following is the product of the reaction below?
$ZnS(s) + O_2 \longrightarrow$
A) SO B) SO_2 C) SO_3 D) $ZnSO_3$ E) $ZnSO_2$

12. Which element is similar to Cd according to its chemical properties?
A) Zn B) Na C) Ca D) Ba E) Sn

13. Cd was discovered in which country?
A) United States B) England C) Germany D) France E)Turkey

14. Which of the following elements gives extra color to emeralds?
A) Al B) Ti C) Fe D) Cr E) Li

15. Which of the following is a product of the reaction below?
$Cr_2O_3 + Al \longrightarrow$
A) Al_2O_3 B) Cr_2O_7 C)CrO_4 D) $Al_2(CrO_4)_3$ E) AlO

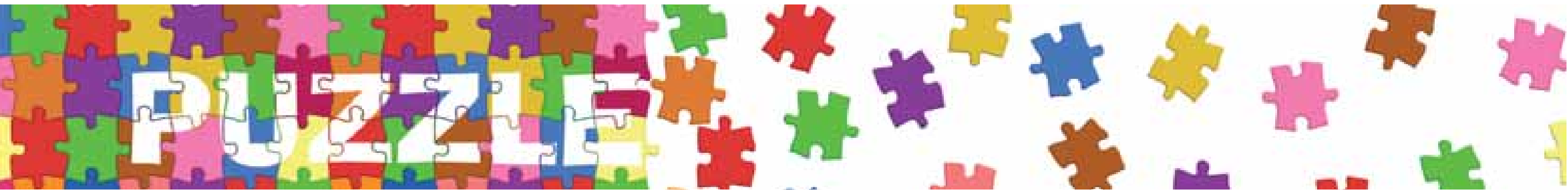
16. Which of the following is the greatest oxidation number that Cr may take?
A) 7 B) 5 C) 4 D) 3 E) 6

17. What's the group of Ti in the periodic table?
A) 1A B) 1B C) 4A D) 3B E) 4B

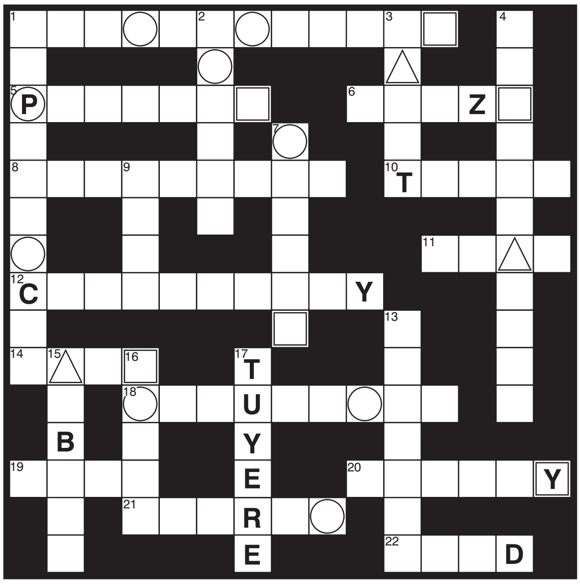
18. Which of the following is a product of the reaction below?
$TiO_2 + 3C + 4Cl_2 \longrightarrow$
A) $TiCl_3$ B) $TiCl_4$ C) Ti D) CO_3 E) $TiCO_3$

19. Which of the following elements is very important for the aircraft industry?
A) Cs B) Ca C) Ti D) Au E) Hg

20. What is the color of titanium dioxide, commonly used in the food industry?
A) Black B) Brown C) Yellow D) White E) Orange



Answer the clues as normal. The letters in the triangles, circles and squares will spell out three more words to do with the subject, if read in order. There are Extra Clues for these words.



CLUES ACROSS

1. Used in the production of iron. (5,7)
5. Most of this, produced in a 1 Across, is converted into steel (3,4)
6. Made of steel, this surrounds the cathodes in a Downs cell. (5)
8. Nitric acid is manufactured when this happens to ammonia. (9)
10. An industrial site should be far away from housing if the process uses such substances. (5)
11. The Contact process produces sulphuric ________ (4)
12. This of production, is all-important if an industrial process is to be economic. (10)
14. Basic ________, a by-product of iron production, is used as a fertiliser. (4)
18. This process is carried out on nitrogen in the production of ammonia. (9)
19. One of the reactants initially loaded in the blast furnace. (4)
20. The catalyst involved in the large-scale production of hydrogen gas has been ________ divided. (6)
21. This acid is produced by the Ostwald process. (6)
22. ________ communication links are needed if an industrial site is to be economic to use. (4)

CLUES DOWN

1. When sodium is extracted using a Downs cell, chlorine is the only one of these. (10)
2. Slag does this on the surface of molten iron in a blast furnace. (6)
3. Many process industries are positioned near to this, as many raw materials have to be imported. (5)
4. Sodium hydrogen carbonate forms a ________ at the bottom of the carbonator used in the Solvay process. (11)
7. This is blown in at the base of a blast furnace. (3,3)
9. The sulphur dioxide used in the Contact process is passed through electrostatic precipitators to remove this. (4)
13. Water may be needed for this purpose, which will affect the siting of industrial plant. (7)
15. People supply this. (6)
16. The colour of the gas which is produced at the anode of a Downs cell, used in the industrial extraction of sodium. (5)
17. 7 Down is blown through this. (6)

Extra Clues

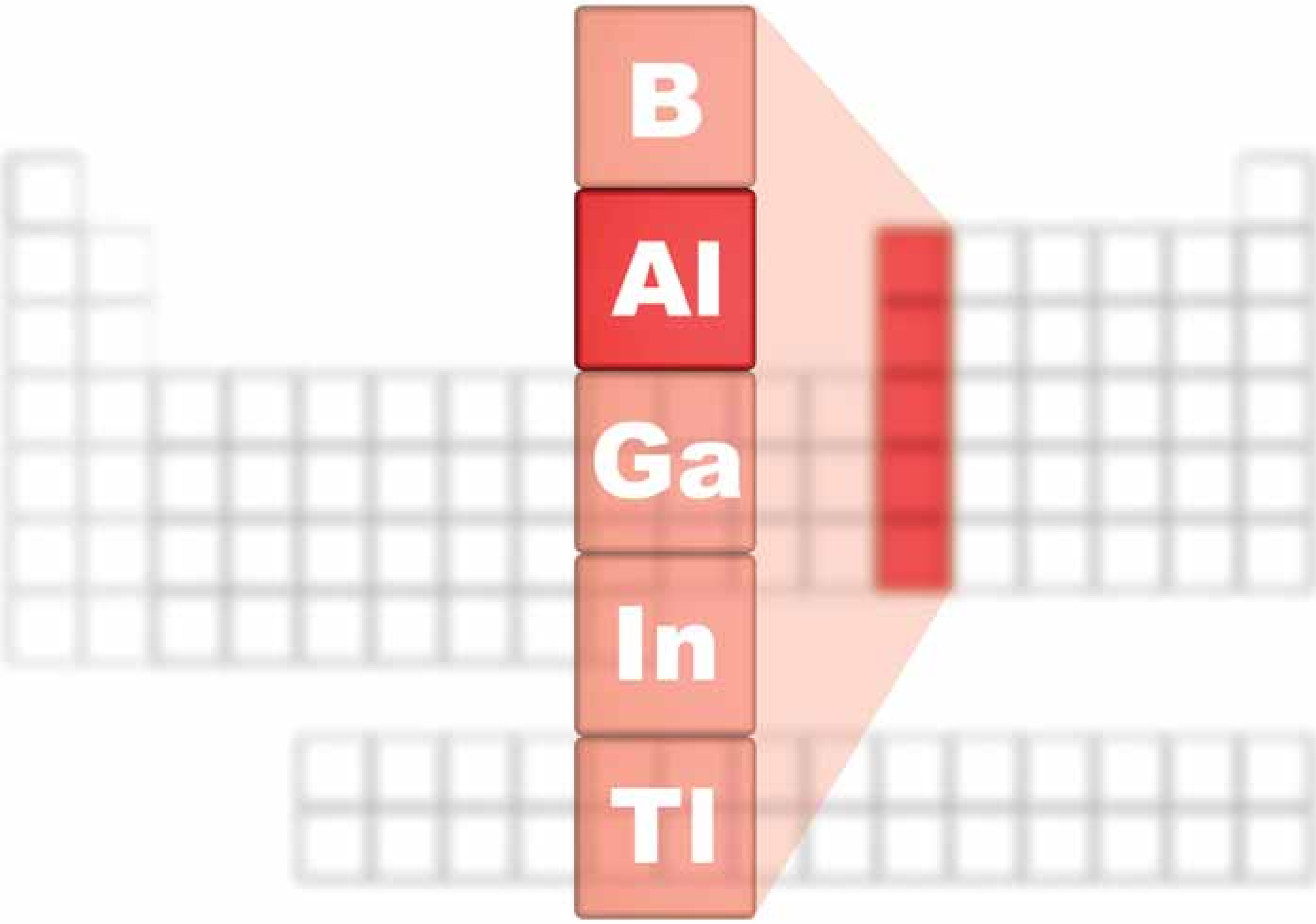
○ This acid is regarded as the most important product of the chemical industry. (9)

□ Industry tries to use the minimum amount of this expensive commodity. (6)

△ In its crude form, this liquid is a complex mixture of hydrocarbons. (3)

NOTES

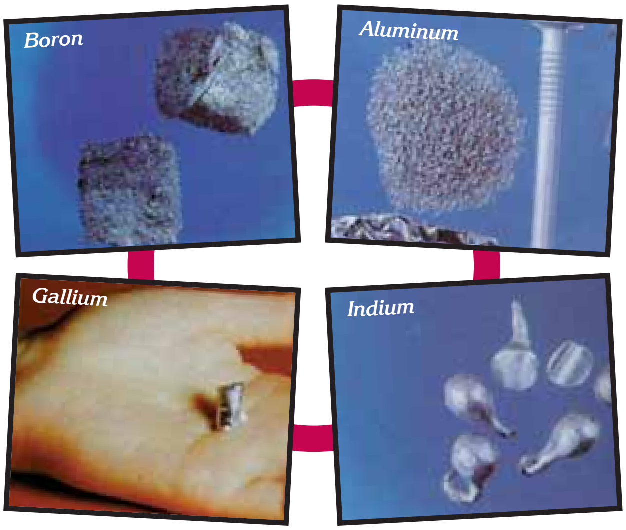
METALS



ALUMINUM

INTRODUCTION

The group 3A elements in the periodic table are boron (B), aluminum (Al), gallium (Ga), indium (In) and thallium (Tl). ***This group is also known as earth metals***.



All of the group 3A elements, except boron, are metals. From bottom to top within a group in the periodic table, the atomic radius decreases. In relation to this trend, as metallic character diminishes, non–metallic character becomes dominant. As we go from thallium to boron, the metallic character of the elements decreases. Hence the element boron, the first member of this group, possesses both metallic and nonmetallic properties to some extent and is called a metalloid, whereas the other members are called metals.

The electron configurations of the group 3A elements end with ns^2np^1. The most important oxidation state of these elements is +3. Moreover, gallium, indium and thallium can take +1 charge by giving up only one of their electrons in the p orbital.

The elements gallium, indium and thallium are found in nature in trace amounts, and are so soft that they can easily be cut with a knife. Gallium melts at 29.7°C, and boils at the very high temperature of 2403°C. Because of this great difference between the melting and boiling

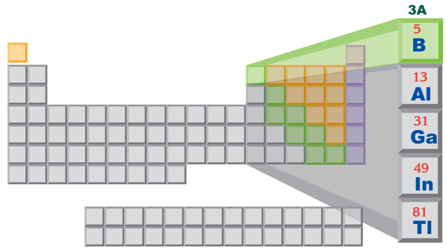


Figure 1 Earth metals, group 3A elements.

points, durable thermometers are which able to withstand high temperatures are designed using boron or gallium.

The element indium has a silvery–white appearance and is used in the manufacture of special mirrors for some optical devices. The element thallium, with properties similar to its neighbor lead, is a brilliant bluish–white metal. Since it is readily oxidized in air, it should be stored under oil.

Among the group 3A elements, boron and aluminum are widely used, important elements in industry. In industrial applications, boron is used in the form of its compounds, whereas aluminum is used as a metal. That's why, in this section, boron and aluminum will be examined in more detail than the others.

Name	Aluminum
Symbol	Al
Atomic number	13
Atomic mass	26.9
Electron configuration	$[Ne]3s^23p^1$
Melting point (°C)	660.3
Boiling point (°C)	2519
Density (g/cm³)	2.7
I^{st} Ionization energy (kj/mol)	577.5
Atomic radius (pm)	118
Common oxidation numbers	+3
Color	silvery
Physical state at 25°C	solid
Origin and meaning of name	*alumen* - alum
Earth's crust abundance (%)	8.2

Table 1 Some properties of Aluminum

1. ALUMINUM

Aluminum was first isolated from the mineral alum in 1827 by the German chemist F. Woehler. The name aluminum comes from the word "alumen" which is the Latin equivalent of alum, meaning bitter taste.

Aluminum, the second member of group 3A, has atomic number 13. Its electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^1$, so it takes +3 oxidation state in its compounds to have the same electron configuration as neon.

$_{13}Al: 1s^2 \; 2s^2 \; 2p^6 \; 3s^2 \; 3p^1$

⇅ ⇅ ⇅⇅⇅ ⇅ ↑○○

Aluminum is a silvery white element with metallic luster. It melts at 660.3°C, boils at 1519°C, and has a density of 2.7 g/cm^3. It is a soft metal, but its hardness can be increased when it is alloyed with elements such as Cu, Mg, Mn and Si.

Aluminum has lots of different uses because of its ductility and malleability. It can easily be hammered into wire and plate, even into sheets. Hence, it is frequently used for packaging food. Because it is an excellent conductor of electricity, and it is lighter and cheaper than copper, it is used in high–voltage lines.

1.1 OCCURRENCE

Aluminum, after oxygen and silicon, is the third most abundant element in the earth's crust. This means that aluminum is the most abundant metal in the earth's crust.

Aluminum cannot be found in its natural state as it has a great affinity for oxygen. That's why aluminum is found in nature in the form of oxides of aluminum. The main aluminum ores are feldspar ($K_2Al_2Si_6O_{16}$), kaolinite ($Al_2Si_2O_7 \cdot 2H_2O$), corundum (Al_2O_3), cryolite (Na_3AlF_6) and bauxite ($Al_2O_3 \cdot 2H_2O$) (Figure 2).

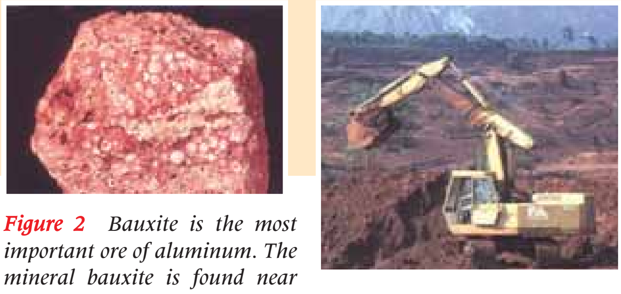


Figure 2 Bauxite is the most important ore of aluminum. The mineral bauxite is found near the surface of the earth. It contains only 50–60% Al_2O_3, together with SiO_2, Fe_2O_3, and a small amount of oxides of some transition elements such as titanium, zirconium and vanadium.

Corundum, an aluminum oxide (Al_2O_3), is a highly stable, hard compound that is naturally transparent. However, some corundum may naturally mix with chrome and acquire red crystalline structures. These crystals are the precious stones we know as rubies, and are used in jewelry. Some corundum may naturally mix with the element cobalt and form a precious stone called sapphire, which has a sky-blue color. All these precious stones, like rubies and sapphires, can also be synthetically produced.

Ruby

Sapphire

Preparation

The industrial production of aluminum includes two successive steps.

1. In the first step, pure Al_2O_3 is obtained from the mineral bauxite. Bauxite contains only 50–60% Al_2O_3, together with SiO_2, Fe_2O_3 and a small amount of oxides of some transition elements such as titanium, zirconium and vanadium.

Bauxite can be purified by the Baeyer's process. In this process, bauxite is powdered and roasted at about 700 °C. It is then mixed with NaOH solution under high pressure at a temperature of about 150–170°C. Hence, the amphoteric Al_2O_3 is turned into $NaAlO_2$. The other oxides, such as Fe_2O_3 and SiO_2, are eliminated from the reaction medium in the form of red mud when the solution is filtered.

The $NaAlO_2$ compound obtained is then dissolved in water. When a solid $Al(OH)_3$ is added to this water solution, all of the Al is precipitated in the form of $Al(OH)_3$.

$$Al_2O_3 + 2NaOH \longrightarrow 2NaAlO_2 + H_2O$$

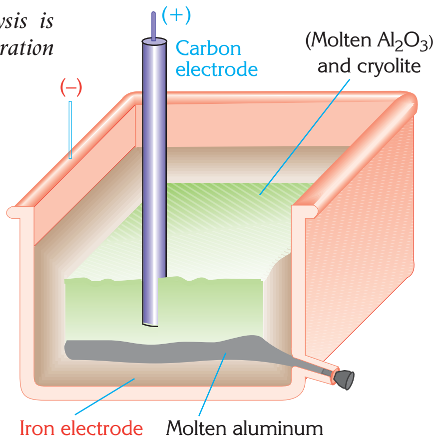
$$NaAlO_2 + 2H_2O \longrightarrow Al(OH)_3\downarrow + NaOH$$

When the $Al(OH)_3$ is heated, the compound Al_2O_3 is produced by loss of water.

$$2Al(OH)_3 \xrightarrow{1200°C} Al_2O_3 + 3H_2O$$

The second step in the preparation of aluminum is the electrolysis of molten Al_2O_3 (Figure 3). However, in order to reduce the melting point of Al_2O_3, some amount of cryolite (Na_3AlF_6) is added to the reaction medium. Thus, the melting point of Al_2O_3 is reduced to about 1000°C from 2050°C.

Figure 3 *Electrolysis is used in the preparation of aluminum*



The container used in electrolysis is made of steel, and its inner surface is coated with carbon, which serves as the cathode. The carbon rods immersed in the mixture of bauxite–cryolite serve as the anode. During electrolysis the density of aluminum causes it to deposit at the bottom of the cell (cathode) where it is tapped off. The corresponding chemical equations of this process can be shown as follows:

Cathode: 4/ $Al^{+3} + 3e^- \longrightarrow Al(l)$

Anode : 3/ $C(s) + 2O^{-2} \longrightarrow CO_2(g) + 4e^-$

Net reaction;

$$4Al^{+3} + 6\,O^{-2} + 3C(s) \longrightarrow 4Al(s) + 3CO_2(g)$$

The purity of the aluminum obtained is between 99.5 - 99.9%. This process was first developed by Charles Martin Hall in 1886 when he was a student at Oberlin College. With this method, aluminum has been cheaply and easily produced to this day.

An aluminum factory

2. A cheaper method has been developed. In this method, anhydrous bauxite is turned into $AlCl_3$ by reacting with chlorine gas in a carbonated medium.

$$2Al_2O_3(s) + 3C(s) + 6Cl_2(g) \rightarrow 4AlCl_3(l) + 3CO_2(g)$$

The molten $AlCl_3$ is then electrolyzed.

$$2AlCl_3(s) \xrightarrow{electrolysis} 2Al(s) + 3Cl_2(g)$$

This process is about 70% more economic than Martin Hall's process.

1.2 CHEMICAL PROPERTIES

Aluminum is an amphoteric metal, so it can react with both acids and bases. In addition it reacts with some nonmetals and reduces some metal oxides.

Reactions

1. Aluminum reacts with diluted HCl and H_2SO_4 solutions.

   $2Al\ (s) + 6HCl\ (dil.) \longrightarrow 2AlCl_3(s) + 3H_2(g)$ (Figure 4).

   $2Al\ (s) + 3H_2SO_4\ (dil.) \longrightarrow Al_2(SO_4)_3(s) + 3H_2(g)$

   Even though it is an active metal, because of the formation of a protective oxide layer on its surface, aluminum becomes very passive in reacting to HNO_3 in a cold medium, and does not react with water.

   Even though HNO_3 does not affect aluminum at room temperature, when its dilute solution is heated it reacts with aluminum.

   $8Al + 30HNO_3(dil.) \xrightarrow{heat} 8Al(NO_3)_3 + 3N_2O\uparrow + 15H_2O$

   Aluminum also can be affected by concentrated HNO_3 and H_2SO_4 only when they are hot.

   $Al + 6HNO_3(conc.) \xrightarrow{heat} Al(NO_3)_3 + 3NO_2\uparrow + 3H_2O$

   $2Al\ (s) + 6H_2SO_4\ (conc.) \xrightarrow{heat} Al_2(SO_4)_3\ (s) + 3SO_2(g) + 6H_2O(g)$

2. Since aluminum is an amphoteric metal it also reacts with strong bases and liberates H_2 gas.

   $2Al + 2NaOH + 2H_2O \xrightarrow{heat} 2NaAlO_2 + 3H_2\uparrow$

   $2Al + 6NaOH + 6H_2O \longrightarrow 2Na_3[Al(OH)_6] + 3H_2\uparrow$ (Figure 5)

   Both the oxide and hydroxide of aluminum (Al_2O_3 and $Al(OH)_3$) are amphoteric compounds.

3. If mercury salts contact the surface of aluminum metal, elemental mercury and aluminum salts are produced.

   $2Al(s) + 3HgCl_2(l) \longrightarrow 2AlCl_3(s) + 3Hg(s)$

   If aluminum is mixed with the mercury formed, an amalgam is obtained. This amalgam reacts with water at room temperature and releases H_2 gas, since the mercury in the amalgam prevents the formation of an oxide layer on the surface of the aluminum.

   $2Al(s) + 6H_2O \longrightarrow 2Al(OH)_3 + 3H_2\uparrow$

Figure 4 Reaction of Al with HCl acid.

$2Al(s) + 6HCl(aq) \longrightarrow 2AlCl_3(s) + 3H_2(g)$

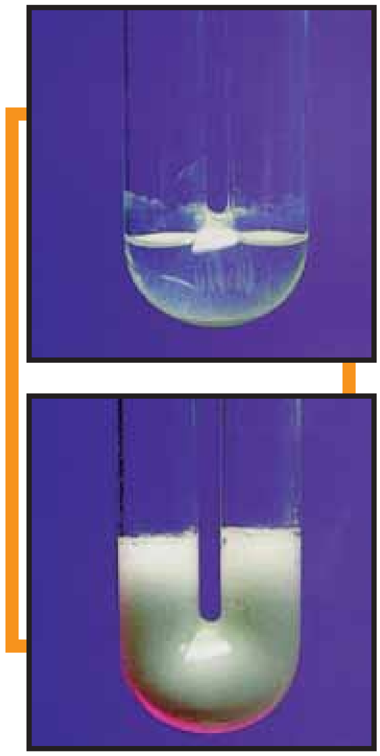


Figure 5 Reaction of aluminum with sodium hydroxide.

4. Aluminum reacts with nonmetals and several compounds directly at high temperature.

$$4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s) + heat$$

$$2Al(s) + 3Br_2(g) \longrightarrow 2AlBr_3(s) + heat \text{ (Figure 6)}$$

$$2Al(s) + 3S(g) \xrightarrow{150\text{–}200°C} Al_2S_3(s)$$

Figure 6 *Reaction between aluminum and bromine*

$$2Al(s) + 3Br_2(g) \longrightarrow 2AlBr_3(s)$$

5. Aluminum has a great affinity for oxygen. This property has good technical applications, such as the reduction of metals from their compounds, e.g. Fe_2O_3 and Cr_2O_3 (Figure 7).

Figure 7 *Reduction of iron from its ore with aluminum.*

$$Fe_2O_3 + 2Al \longrightarrow Al_2O_3 + 2Fe$$

$$Fe_2O_3 + 2Al \longrightarrow Al_2O_3 + 2Fe$$

$$Cr_2O_3 + 2Al \longrightarrow Al_2O_3 + 2Cr$$

During the course of these reactions, the temperature rises to about 2000 °C, and the molten metal produced at this temperature can be easily isolated from the reaction medium.

1.3 COMPOUNDS

1. Aluminum oxide, Al_2O_3

The commercial name of Al_2O_3 is alumina. It is naturally found as the mineral bauxite. Al_2O_3 is a white, hard substance with a melting point of 2045°C. Al_2O_3 is almost insoluble in water. As mentioned before, alumina shows amphoteric properties like aluminum, and reacts with acids and bases.

Aluminum oxide, Al_2O_3 is a white solid known as alumina.

$$Al_2O_3(s) + 6HCl(conc.) \xrightarrow{heat} 2AlCl_3(s) + 3H_2O(l)$$

$$Al_2O_3(s) + 2NaOH(conc.) \xrightarrow{900 - 1100°C} 2NaAlO_2(s) + H_2O(l)$$

$$Al_2O_3(s) + 2NaOH(hot, conc.) + 3H_2O \longrightarrow 2Na[Al(OH)_4]$$

> *$NaAlO_2$ can be hydrated by the removal of one or two water molecules and can be formulated as:*
>
> *a) $NaAlO_2 . H_2O$ or NaH_2AlO_3*
>
> *b) $NaAlO_2 . 2H_2O$ or $Na[Al(OH)_4]$*

Preparation

Since aluminum has a great affinity for oxygen, when heated strongly in air, it produces Al_2O_3 with evolution of a huge amount of heat.

$$4Al(s) + 3O_2(g) \xrightarrow{heat} 2Al_2O_3(s) \quad \Delta H = -400 \text{ kcal/mol}$$

2. Aluminum hydroxide, $Al(OH)_3$

Aluminum hydroxide is a gray precipitate produced from the reaction of aluminum salts with bases.

$$AlCl_3(aq) + 3NaOH(aq) \longrightarrow Al(OH)_3(s) + 3NaCl(aq)$$

This is the method for detecting Al^{3+} ions in analytic chemistry.

$$Al^{3+} + 3OH^- \longrightarrow Al(OH)_3\downarrow$$

If the medium contains more base, it causes $Al(OH)_3$ to dissolve.

$$Al(OH)_3(s) + OH^-(aq) \longrightarrow [Al(OH)_4]^-(aq)$$

Like Al_2O_3, $Al(OH)_3$ is an amphoteric substance, so it reacts with bases and acids.

$$Al(OH)_3 + 3HCl(dil.) \longrightarrow AlCl_3 + 3H_2O$$

$$Al(OH)_3(s) + NaOH(conc.) \longrightarrow Na[Al(OH)_4](s)$$

Even though $Al(OH)_3$ gives such reactions with acids and bases, it is not soluble in water.

> *If we put basic foods in aluminum pots, toxic aluminate forms.*
>
> $2Al + 2OH^- + 6H_2O \longrightarrow 2Al(OH)_4^- + 3H_2$

3. Aluminum sulfate, $Al_2(SO_4)_3$

Aluminum sulfate forms in nature as an important series of alums, $MAl(SO_4)_2 \cdot 12H_2O$, where M may be almost any univalent cation such as Na^+, K^+, NH_4^+ and Ag^+.

Aluminum sulfate is soluble in water like the nitrate and halide salts of aluminum. In industry, aluminum sulfate is obtained by the reaction of Al_2O_3 with hot, concentrated sulfuric acid.

$$Al_2O_3(l) + 3H_2SO_4(conc.) \xrightarrow{heat} Al_2(SO_4)_3(aq) + 3H_2O(l)$$

In the laboratory it is obtained by the reaction of aluminum with the sulfate salts of less reactive metals.

$$2Al(s) + 3CuSO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + 3Cu(s)$$

(Figure 8)

$Al_2(SO_4)_3$ is used in the treatment of sea water and other reservoirs (in order to precipitate impurities), in the paper industry (to give strength to the paper), in the production of water–proof goods, and in the dye industry.

Figure 8 Reaction of Al with copper sulfate.

Name	Percentage Composition	Properties	Use
Duralumin	Al : over 90% Cu : about 4% Mg : 0.5 - 1% Mn : Less than 1%	hard, durable, light	aircraft plating automobile bodies propeller construction
Magnalium	Mg : about 5% Al : remainder	lighter and more workable than aluminum hard, durable	production of engineering materials metal mirrors scientific instruments construction of plane body light electroplating
Aluminum Bronze	Al : about 10% Fe : about 3% Ni : about 3% Mn : about 2% Cu : about 80%	high strength corrosion resistance higher than brass	bearing applications marine applications like propellers production of machine parts

A vehicle manufactured with aluminum alloy

Table 2 The most important alloys of aluminum.

Metals

Uses

Aluminum is used as the main element of light alloys with high electrical conductivity, durability, and high resistance to corrosion, such as duralumin (Al, Mg, Cu, Mn), magnalium (Mg, Al) and aluminum bronze (Al, Cu). These alloys are used in the manufacture of planes, ships, engines, submarines and automobiles, and also play important roles in aerospace technology.

Aluminum is used at home because of its low cost, appearance, lightness and good conductivity of heat and electricity. Aluminum is also used for packaging foods.

Since aluminum is a light metal with high electrical conductivity, it is preferred over copper for high voltage electrical lines.

Aluminum powders are used in camera flashes and in the dye industry. Since these powders can be easily burnt in air causing an explosion, they are used in the aluminothermic process and in the production of thermite for napalm bombs.

SUPPLEMENTARY QUESTIONS

1. Aluminum carbide and hydrochloric acid reaction is like in the equation below.

$Al_4C_3 + HCl(g) \longrightarrow AlCl_3 + CH_4(g)$

According to the reaction, how many grams of methane (CH_4) gas can be produced from 0,5 mol aluminum carbide and 8 moles HCl at STP?

2. Write balanced equations for the reactions of Al metal with dilute HCl and NaOH solutions.

3. How many moles of aluminum will be needed for the reaction of HCl to produce the same amount of hydrogen from decomposition of 6,8 g NH_3?

4. How many liters of H_2gas can be obtained from the reaction of 5,4 g Al and 20 g concentrated NaOH at STP.

5. How many grams of $Al_2(SO_4)_3$ will be needed to prepare a 500 ml solution which has $[SO_4^{-2}]$ = 0,18 M?

6. Complete the following reactions.

a. $Al(OH)_3(s) + NaOH(aq) \longrightarrow$

b. $Al_2O_3(s) + H_2SO_4(aq) \longrightarrow$

c. $2Al_2O_3(s) + 3Cl_2(g) + 3C(s) \longrightarrow$

d. $Al_2O_3(s) + H_2(g) + NaOH \longrightarrow$

7. Where do we use aluminum in our daily life?

8. What are the important aluminum ores?

9. Write two reactions to show amphoteric property of aluminum oxide, Al_2O_3.

10. Write electronic configuration of $_{13}Al$.

11. Explain the thermite process and write the equation for this reaction.

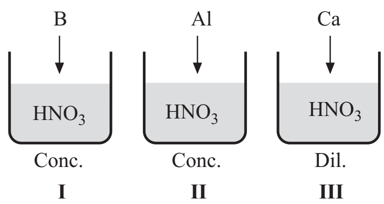
12. How can you separate B - Al mixture by using water? Write the possible reactions.

13. The formulas of $B(OH)_3$ and $Al(OH)_3$ are comparable although their chemical properties are quite different. Explain the reason of that.

14. 27,2 g of the mixture Al_4C_3 and CaC_2 reacted with HCl solution. Obtained gases are burned in excess oxygen and 22,4 L CO_2 gas is obtained at STP. Find the mol ratio between Al_4C_3 and CaC_2 in mixture.

15. These are three test tubes which are filled with NaCl, $Ca(NO_3)_2$ and $Al(NO_3)_3$ respectively. By using a base solution how can be found the given chemicals correctly? Explain.

16.



When the elements are put into showed hot solutions, in which one(s) gas evolution is/are observed.

Write chemical equations for possible reactions.

17. Perform the following conversions.

a. $Al \rightarrow Al_2O_3 \rightarrow AlCl_3 \rightarrow Al \rightarrow Al_4C_3 \rightarrow Al(OH)_3 \rightarrow K[Al(OH)_4]$

b. $Al \rightarrow Al_2(SO_4)_3 \rightarrow Al(OH)_3 \rightarrow Al(NO_3)_3 \rightarrow Al_2O_3 \rightarrow NaAlO_2$

MULTIPLE CHOICE QUESTIONS

1. Which one is the electron configuration of aluminum?

A) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$ B) $1s^2 2s^2 p^6 3s^2 3p^1$ C) $1s^2 s^2 2p^3$ D) $1s^2 2s^2 2p^1$ E) $1s^2 2s^2 2p^6 3s^2$

2. Which one(s) give(s) reaction with aluminum hydroxide?

A) N_2 B) NaOH C) Na_2SO_4 D) H_2O E) Ag

3. Which substance is produced by the reaction of Al_2O_3 and NaOH?

A) Na_2O B) AlO_3 C) $NaAlO_2$ D) H_2 E) $Al(OH)_4$

4. Which one of he following metals is/more active than aluminum?

A) Cr B) Cu C) Ag D) Ca E) Fe

5. Which one(s) give reaction readily with aluminum oxide Al_2O_3?

A) H_2O B) H_2SO_4(dil.) C) N_2 D) NaOH E) Pt

6. Which one doesn't react with aluminum?

A) HCl B) S C) P D) KOH E) All react

7. Which one of the following is not an earth metal?

A) Al B) Ga C) In D) Tl E) Si

8. Which one of the following metals can not be handed for a long time as a solid?

A) Al B) Ga C) Fe D) Zn E) Pb

9. What is the place of the Al among the metals exist in the earth's crust by abundance?

A) 1 B) 2 C) 3 D) 4 E) 5

10. Which one of the following is the first member of the group 3A?

A) Tl B) In C) Ga D) Al E) B

11. Which one of the following does not exist in the bauxite mineral?

A) CaO B) Al_2O_3 C) SiO_2 D) Fe_2O_3 E) H_2O

12. Which one of the following is not amphoteric metal?

A) Al B) Zn C) Cr D) Br E) Sn

13. What is the name of the compound, Na_3AlO_3?

A) Aluminum sodium oxide
B) Sodium aluminate
C) Sodium aluminum oxide
D) Aluminum (III) Sodium oxide
E) Complex oxide

14. Which element is produced at the end of the termite reaction?

A) Al B) K C) Fe D) Li E) F

15. Which one of the following is the most important compound of aluminum used in commercial life?

A) $Al_2(SO_4)_3$ B) Al_2O_3 C) $Al_2(SO_3)_3$ D) $AlCl_3$ E) $Al_2(CO_3)_3$

16. The first ionization energy of 3A group elements is greater than the energy of 2A group elements. Why?

A) They have smaller atomic radius.
B) They have bigger atomic radius.
C) They have less electron affinity value.
D) Electrons from p-orbital need less energy than the electrons from s orbital.
E) They can be easily ionised.

17. Although copper is a better conductor of electricity than aluminum, Al is used in cables.Why?

A) Because of durability and resistivity toward the corrosion
B) Because of number of electrons in the left energy level
C) Because of abundance in nature
D) Because it's lighter and cheaper
E) None of them

18. Why is the aluminum metal not found in elemental form in the earth's crust?

A) It has a tendency for oxygen
B) Because it is an active metal
C) It has amphoteric property
D) It is unstable
E) Because it is not an active metal

19. Which of the following do(es) not give reaction with aluminum?

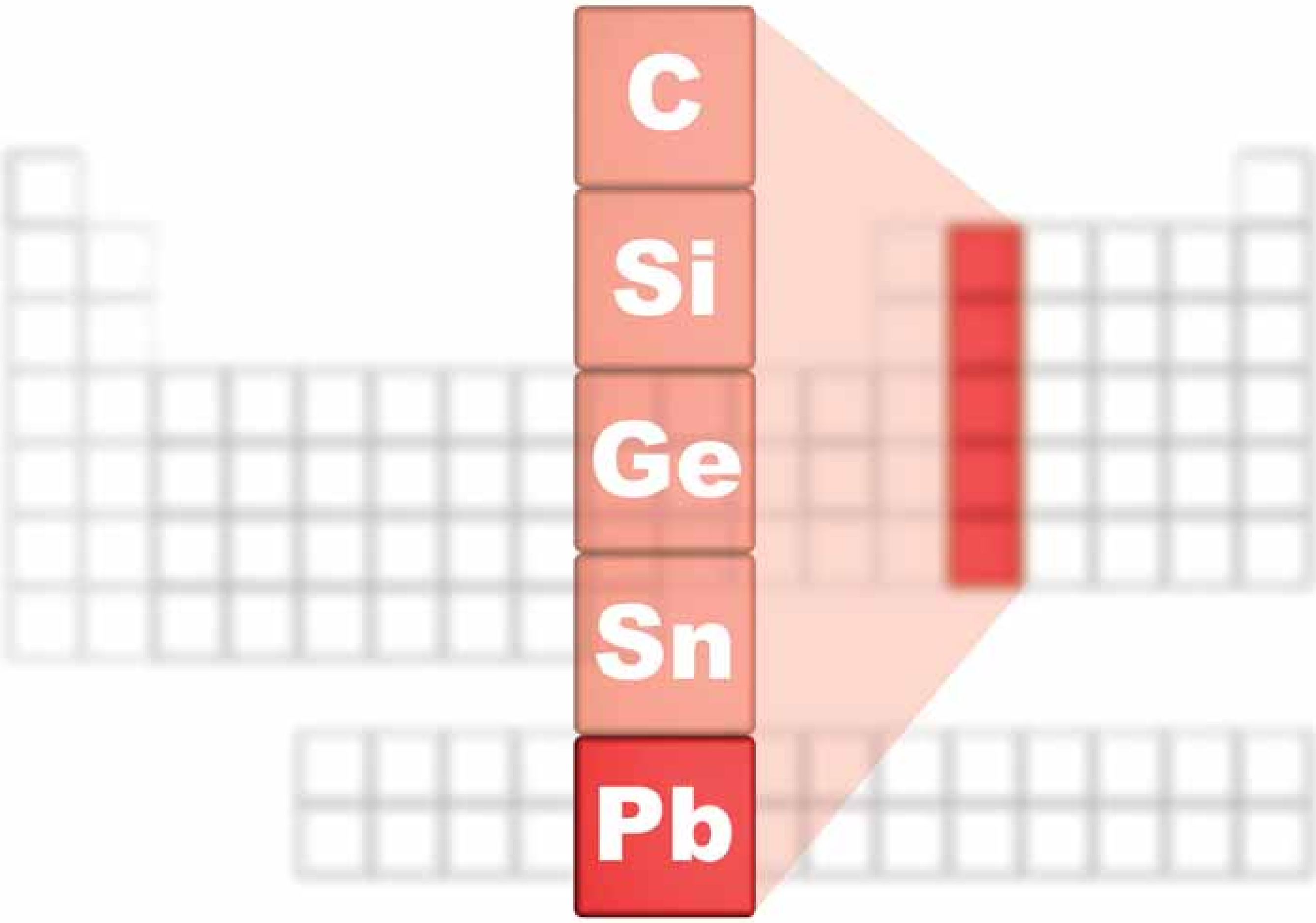
I. HCl II. H_2SO_4 III. NaOH IV. CO

A) I only B) I and II C) I, III and IV D) IV only E) I and II

20. Which one of the following has covalent bond in its compounds?

A) Al B) Ba C) B D) Pt E) Ag

METALS

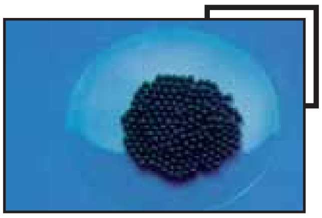


LEAD

INTRODUCTION

In group 4A of the periodic table carbon (C), silicon (Si), germanium (Ge), tin (Sn) and lead (Pb) are placed. ***Because the first member of the group is carbon, group 4A is called the carbon group or carbon family.***

Germanium



Lead

Tin

Silicon

Name	Lead
Symbol	Pb
Atomic number	82
Atomic mass	207.2
Electron configuration	$[Xe]4f^{14}5d^{10}6s^{2}6p^{2}$
Melting point (°C)	327.5
Boiling point (°C)	1749
Density (g/cm³)	11.34
Ist Ionization energy (kj/mol)	715.6
Atomic radius (pm)	147
Common oxidation numbers	+2, +4
Color	bluish white
Physical state at 25°C	solid
Origin and meaning of name	*plumbum* - liquid silver
Earth's crust abundance (%)	10^{-3}

Table 1 Some properties of lead

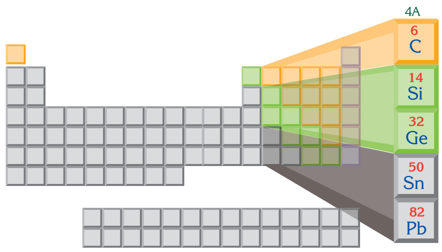


Figure 1 Place of group 4A elements in periodic table

Carbon is one of the most well-known elements in the world. Carbon, the lightest member of the group, is a nonmetal. The next two members, silicon and germanium, are metalloids. Silicon is mostly nonmetal and germanium is mostly metal in properties. Tin, except only one nonmetal isotope (grey tin), is a metal with all properties. Pb is the only metal in the group. These differences in the group are related to the atomic radii of the elements. In the group, the atomic radius increases from top to bottom, and, for that reason, the tendency to give electrons, or metallic property of the elements, increases.

Electron configurations of group 4A elements end with ns^2np^2, and they have –4, +2 and +4 oxidation states in their compounds.

Carbon and silicon make covalent, tin and lead make ionic, and germanium makes both ionic and covalent bonds.

In group 4A, as atomic number and atomic mass increase, melting and boiling points decrease and densities increase.

Graphite, an allotrope of carbon, silicon, and germanium are poor conductors of electricity at room conditions. With increasing temperature, their conductivities also increase. The other members of the group, tin and lead, are good conductors of electricity. But an increase in temperature causes a decrease in the conductivity of other metals.

Now, let's study lead in detail.

1. LEAD

Lead comes from the Latin word "plumbus", which means heavy. Lead is the last member of the carbon family and shows metallic properties. Its atomic number is 82 and its electron configuration is $[_{54}Xe]\ 4f^{14}\ 6s^2\ 5d^{10}\ 6p^2$.

Lead is a grey colored, shiny, soft metal. It can be cut easily. It is malleable but not ductile. Lead has the lowest melting point of group 4A elements (327.5°C). Because of its high density (11.34 g/cm^3), it is accepted as a heavy metal.

Freshly cut lead has a shiny surface but it quickly oxidizes with oxygen in the air and becomes dull. The oxidized layer on the surface of lead protects the rest of the metal from oxidation.

Lead and lead compounds are poisonous because heavy metals are deposited easily in the body. Lead prevents enzymes from catalyzing some important reactions.

1.1 OCCURRENCE

Lead is found in trace amounts in the earth's crust (0.001%) as its minerals, such as galena (PbS), cerussite ($PbCO_3$) and anglesite ($PbSO_4$).

Preparation

The most commonly used source of lead is galena (PbS). In the preparation of lead, powdered galena is roasted in a blast furnace by "partial oxidation" method.

$$2PbS(s) + 3O_2(g) \xrightarrow{1200°C} 2PbO(s) + 2SO_2(g) \text{ or}$$

$$PbS(s) + 2O_2(g) \xrightarrow{300\text{-}400°C} PbSO_4(s)$$

After this step, the temperature is increased to 850°C. The remaining galena (PbS) is treated with PbO and $PbSO_4$ to produce molten lead.

$$PbS(k) + 2PbO(s) \longrightarrow 3Pb(l) + SO_2(g) \text{ or}$$

$$PbS(k) + PbSO_4(s) \longrightarrow 2Pb(l) + 2SO_2(g)$$

If necessary, this molten lead can be purified by electrolysis.

> *Heavy metals are the metals with density greater than 5 g/cm^3.*

1.2 CHEMICAL PROPERTIES

Lead takes +2 and +4 oxidation states in its compounds. It is an amphoteric metal. Lead reacts with chalcogens, halogens, strong bases and some acids at suitable conditions. Pb^{2+} ion has a poisonous effect on living organisms.

Reactions

1. Pb forms its halides with halogens by direct reactions.

$$Pb(s) + F_2(g) \xrightarrow{200\text{-}300°C} PbF_2(s)$$

$$Pb(s) + Cl_2(g) \xrightarrow{200\text{-}300°C} PbCl_2(s)$$

2. Lead powder reacts with oxygen at high temperature to give Pb or Pb_3O_4. First PbO is formed. If oxygen is in excess, Pb_3O_4 (red lead) is formed.

$$2Pb + O_2 \xrightarrow{600°C} 2PbO(\text{yellow})$$

$$6PbO + O_2 \xrightarrow{450°C} 2Pb_3O_4(\text{orange-red})$$

Pb_3O_4 is a mixed oxide of PbO_2 and 2PbO and can be formulated as $PbO_2 \cdot 2PbO$.

3. When O_2 and CO_2 are dissolved in water passing through lead pipes, they react with lead to form white $Pb(OH)_2$ and $PbCO_3$ solids.

$2Pb + O_2 + 2H_2O \longrightarrow 2Pb(OH)_2$

$Pb(OH)_2 + CO_2 \longrightarrow PbCO_3 + H_2O$

$2Pb + O_2 + H_2O + CO_2 \longrightarrow PbCO_3 + Pb(OH)_2$ or $Pb_2CO_3(OH)_2(s)$

CO_2 in water changes insoluble $PbCO_3$ into water soluble $Pb(HCO_3)_2$. As a result, lead-covered roofs produces poisonous Pb^{2+} ions when they come in contact with water.

4. Lead easily reacts with sulfur to form black colored lead sulfide.

$$Pb(s) + S(s) \xrightarrow{800\text{-}1200^\circ C} PbS(s)\ (black)$$

5. Lead reacts with dilute nitric acid and concentrated sulfuric and binary acids. It is passive with H_2O, HCl, dilute H_2SO_4 and concentrated HNO_3.

$$3Pb + 8HNO_3(diluted) \longrightarrow 3Pb(NO_3)_2 + 2NO + 4H_2O$$

$$Pb + 2H_2SO_4(conc.) \longrightarrow PbSO_4 + SO_2 + 2H_2O$$

6. Since lead is an amphoteric metal, it reacts with concentrated strong basic solutions.

$$Pb + 2NaOH(conc.) + 2H_2O \longrightarrow Na_2[Pb(OH)_4] + H_2\uparrow$$

Boiler scale building up on the inside of a lead water pipe. If hard water passes through water pipes, the inside of the pipe is covered with $CaCO_3$, $CaSO_4$ and $MgCO_3$. Therefore, lead can not be used in water pipes, thus preventing the transfer of poisonous Pb^{+2} ions to the water.

1.3 COMPOUNDS

Lead has +2 and +4 oxidation states in its compounds. Oxides of lead (PbO, PbO_2, Pb_3O_4), lead sulfide (PbS) (Figure 2) and lead nitrate ($Pb(NO_3)_2$) are commonly used compounds of lead. Now they will be considered in detail.

Figure 2 *Compounds of lead: PbO (yellow), Pb_3O_4 (red), PbS (black)*

1. Lead (II) oxide, PbO

In industry, lead(II) oxide is obtained by the oxidation of lead, Pb, at very high temperature.

$$2Pb(s) + O_2(g) \longrightarrow 2PbO(s)$$

In the laboratory, PbO (yellow) is obtained by the decomposition of $PbCO_3$, $Pb(NO_3)_2$ and Pb_3O_4 compounds when they are heated.

$$2Pb(NO_3)_2(s) \longrightarrow 2PbO(s) + 4NO_2(g) + O_2(g)$$

Powdered PbO is used in glass and ceramics, medicines and in the preparation of glues.

PbO is an amphoteric compound, so it reacts with both acids and bases, but doesn't react with water.

$$PbO(s) + 2HCl(dil.) \longrightarrow PbCl_2(s) + H_2O(g)$$

$$PbO(s) + 2HNO_3(dil.) \longrightarrow Pb(NO_3)_2(aq) + H_2O(l)$$

$$PbO(s) + 2NaOH(aq) \xrightarrow{400^\circ C} Na_2PbO_2(aq) + H_2O(l)$$

2. Lead (IV) oxide, PbO_2

It is obtained by the oxidation of Pb^{2+} in PbO in an oxidizing basic solution.

$$PbO(s) + ClO^-(aq) \longrightarrow PbO_2(s) + Cl^-(aq)$$

PbO_2, a brown colored solid, is an amphoteric oxide with an oxidizing property. It forms Pb^{2+} salts with acids and plumbate (Pb^{4+}) salts with basic oxides and concentrated basic solutions.

$$PbO_2(s) + 4HCl(conc.) \longrightarrow PbCl_2(s) + Cl_2(g) + 2H_2O(l)$$

$$PbO_2(l) + 2H_2S(g)) \longrightarrow PbS(s) + S(s) + 2H_2O(l)$$

$$PbO_2(s) + 2NaOH(conc.) \longrightarrow Na_2PbO_3(aq) + H_2O(l)$$

EXTRA

Why is lead added to gasoline and why is lead-free gasoline used in new cars?

Tetraethyl lead has been used for more than 40 years to improve the combustion characteristics of gasoline. It reduces or eliminates "knocking" (pinging caused by premature ignition) in large, high-performance engines and in smaller, high-compression engines. It provides lubrication to the extremely close-fitting engine parts, where oil has a tendency to wash away or burn off. However, lead will ruin and effectively destroy the catalyst presently used in emission control devices installed in new cars. So lead-free gasoline must be used.

3. Lead (II, IV) oxide, Pb_3O_4

Pb_3O_4, also known as red lead, is formed by heating solid PbO in air to 450°C. If the temperature is increased, it is decomposed into PbO and O_2.

$$6PbO(s) + O_2(g) \xrightarrow{450°C} 2Pb_3O_4(s)$$

$$2Pb_3O_4(s) \xrightarrow{550°C} 6PbO(s) + O_2(g)$$

Pb_3O_4 is also obtained by heating PbO_2 to 350°C under pressure.

$$3PbO_2(s) \xrightarrow{350°C} Pb_3O_4(s) + O_2(g)$$

In fact, Pb_3O_4 is the combination of PbO and PbO_2. That's why it is also shown as $PbO_2 \cdot 2PbO$. It is used in the glass and dye industries.

4. Lead (II) sulfide, PbS

Lead (II) sulfide is a black colored solid obtained by precipitating Pb^{2+} ions from solution with hydrogen sulfide, sodium sulfide or ammonium sulfide.

$$Pb^{2+}(aq) + H_2S(g) \longrightarrow PbS(s) + 2H^+(aq)$$

If Pb^{4+} is used in this reaction, sulfur is seen as a by-product near PbS.

5. Lead (II) nitrate, $Pb(NO_3)_2$

It is the only soluble salt of lead, and is obtained by the reaction of Pb_3O_4 with nitric acid.

$$Pb_3O_4(s) + 4HNO_3(dil.) \longrightarrow 2Pb(NO_3)_2(aq) + PbO_2(s) + 2H_2O(l)$$

In the laboratory, $Pb(NO_3)_2$ is widely used in qualitative analysis.

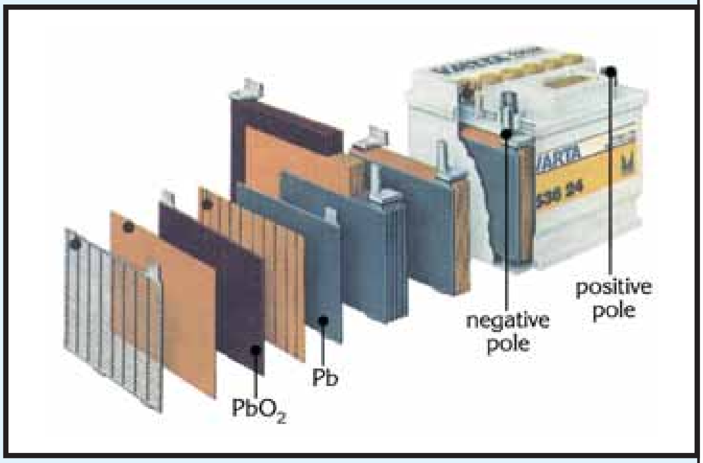
Uses

Lead, is a raw material in the dye industry, and is used in the preparation of Pb_3O_4, which resists corrosion. Alloys of lead are used in fuses and automatic fire extinguishers. An alloy of As-Pb is used in manufacturing pellets.

Lead has a protective property against radiation because of its high density, capability and availability. At the same time, it is not affected by sulfuric acid, so it is used in Pb batteries and in covering underwater cable.

Without the use of lead solders and leaded glass you would not be able to sit safely in front of your PC.

Lead alloys solders give ability to your computer to send electronic data.



Battery

Lead is the glue that binds our electronic world together.

Many of the glass colors themselves are produced with lead.

In space shuttle's lead-alloy solders are used widely.

No other means of connecting transistors, relays and other electronic components is as reliable.

Lead glazes are used to protect the latest generation of electronic microcircuits from atmospheric corrosion.

SUPPLEMENTARY QUESTIONS

1. What is the most important substance of lead? If it is roasted in the blast furnace, which substances can be obtained?

2. Complete the given reactions (if possible) and balance.

a. $Pb + Br_2 \xrightarrow{heat}$

b. $Pb + Se \xrightarrow{heat}$

c. $Pb + O_2 \longrightarrow$

d. $Pb + H_2O \xrightarrow{25°C}$

e. $Pb + CH_3COOH \xrightarrow{+O_2}$

f. $Pb + KOH \xrightarrow{+H_2O}$

3. Research and collect information about the poisonous effect of lead ions.

4. Write the colors of the given compounds of lead.
PbO, PbO_2, Pb_3O_4, PbS, $Pb(NO_3)_2$, $Pb(OH)_2$, $PbSO_4$, $PbCO_3$ and $PbCrO_4$

5. Write the thermal decomposition reaction of $Pb(NO_3)_2$. Which metal nitrates decompose in this way? Give examples.

6. Which of the following reaction(s) occur?

a. $PbO_2 + SO_2 \longrightarrow$

b. $PbO + CO \longrightarrow$

c. $PbO + CO_2 + H_2O \longrightarrow$

d. $PbO_2 \xrightarrow{heat}$

e. Pb (in air) $\xrightarrow{heat}$

f. $Pb^{2+} + Na_2S \longrightarrow$

g. $Pb^{2+} + H_2S \longrightarrow$

h. $Pb^{4+} + (NH_4)_2S \longrightarrow$

i. $PbO + NaOH \longrightarrow$

j. $PbO_2 + NaOH \longrightarrow$

7. Write 5 salts of lead, given in this chapter. Give their colors and solubilities.

8. Predict the products of the given reactions;

a. $Pb_3O_4 + HCl(conc.) \longrightarrow$

b. $Pb_3O_4 + NaOH(conc.) \longrightarrow$

c. $Pb_3O_4 + HNO_3(dil.) \longrightarrow$

d. $Pb_3O_4 \longrightarrow$

9. How can we recycle the lead found in batteries?

10. Which properties of lead allow it to be used in nuclear reactors?

MULTIPLE CHOICE QUESTIONS

1. Which can form ionic bonds?

I. Carbon
II. Silicon
III. Lead

A) I only B) III only C) I and II D) I and III E) I, II and III

2. When X– ions are treated with $Pb(NO_3)_2$ solution, a black precipitation forms. Which of the following may be the X^{2-} ion?

A) Cl^- B) I^- C) NO_2^- D) S^{2-} E) SO_4^{2-}

3. I. It is a heavy metal.
II. It is a transition metal.
III. All its compounds are poisonous.

Which of the above is/are incorrect for lead?

A) I only B) II only C) III only D) I and II E) II and III

4. Which of the following oxidation numbers of lead are given incorrectly?

I. PbO +2
II. PbO_2 +4
III. Pb_3O_4 +4/3

A) I only B) III only C) I and II D) I, II and III E) II and III

5. Which classification(s) is/are correct for oxides of lead?

I. PbO amphoteric oxide
II. PbO_2 basic oxide
III. Pb_3O_4 mixed oxide

A) I only B) I and II C) I and III D) II and III E) I, II and III

6. How many moles of lead salt can be obtained by reaction of 0.5 mol of Pb_3O_4 with sufficient of nitric acid?

(Use the given reactions. First balance the equation, then solve)

$Pb_3O_4(s) + HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + PbO_2(s) + H_2O(l)$

A) 0.25 B) 0.5 C) 1 D) 2 E) 4

7. Which of the following reactions involve a change in oxidation number of a group 4A element from +4 to +2

I. The action of concentrated hydrochloric acid on PbO_2
II. The action of carbon on CO_2
III. The action of heat on SiO_2

A) I only B) II only C) I and II D) II and III E) I, II and III

8. A metal has two oxides. The multiple proportion of oxygen in these oxides equals 3/4.

When oxygen rich oxide is heated, it decomposes to another oxide and oxygen. To obtain 3.36 L oxygen gas at STP, how many moles of oxide must be heated?

A) 0,10 B) 0,15 C) 0,3 D) 0,4 E) 0,6

9. In preparation of lead from galena (PbS), which of the following reactions does not take place?

A) $2PbS + 3O_2 \longrightarrow 2PbO + 2SO_2$
B) $PbS + 2O_2 \longrightarrow PbSO_4$
C) $PbS + 2PbO \longrightarrow 3Pb + SO_2$
D) $PbS + Pb \longrightarrow 2Pb + S$
E) $PbS + PbSO_4 \longrightarrow 2Pb + 2SO_2$

10. During the dissolution of $Pb_2CO_3(OH)_2$ which ion can't form in solution?

A) Pb^{2+} B) CO_3^{2-} C) OH^- D) Pb^{4+} E) HCO_3^-

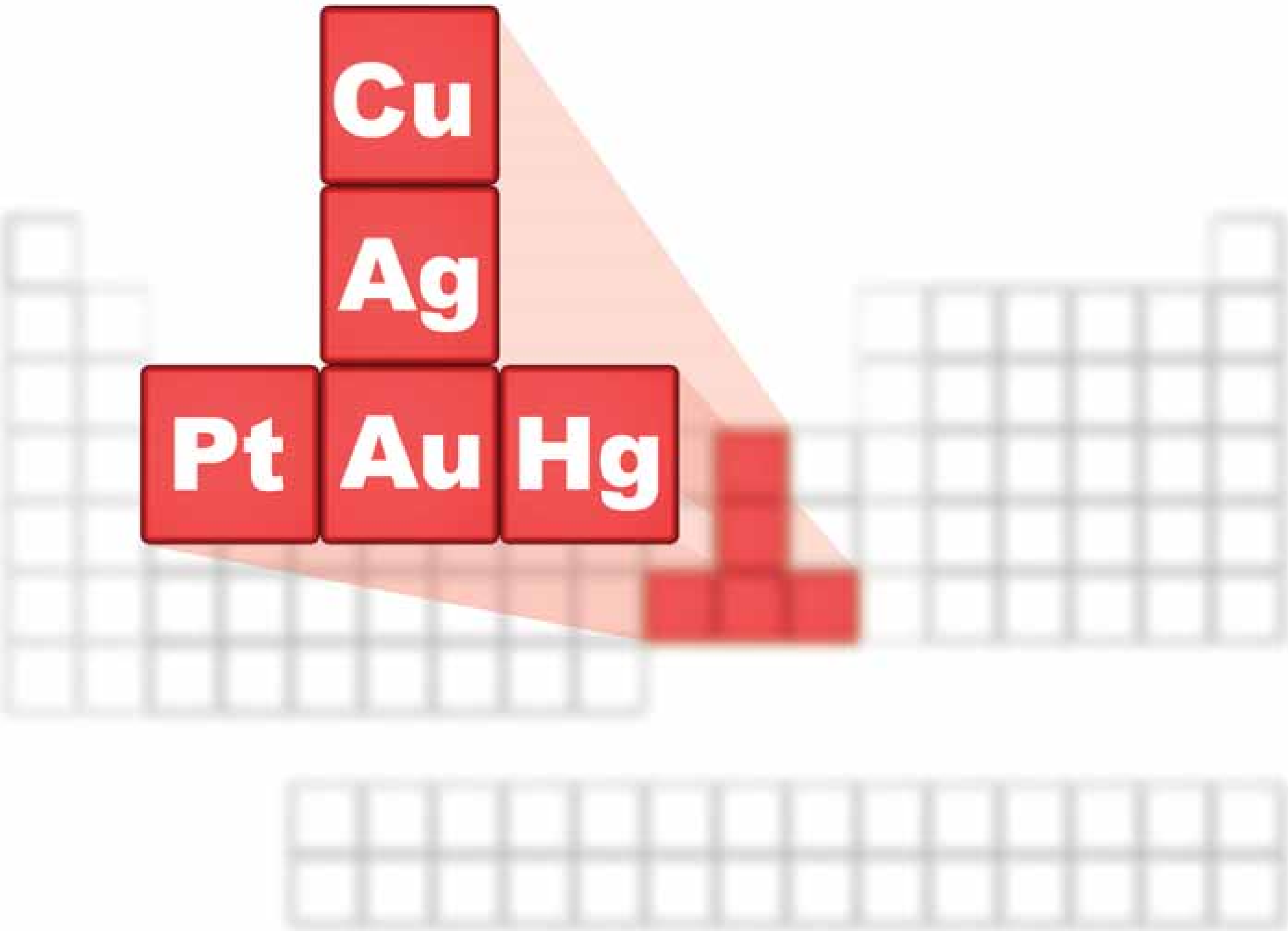
11. I. PbS
II. S
III. H^+

Which of the above can be the product(s) of the given reaction?

$Pb^{4+} + H_2S \longrightarrow$ ....

A) I only B) III only C) I and II D) I and III E) I, II and III

METALS



INERT METALS

INTRODUCTION

Generally metals which are not affected by hydrochloric acid are called inert metals. These metals are less active than hydrogen. Bismuth (Bi), copper (Cu), mercury (Hg), silver (Ag), gold (Au), platinum (Pt), palladium (Pd), osmium (Os), iridium (Ir), ruthenium (Ru) and rhodium (Rh) are inert metals.

Some of the properties of inert metals which are different from other metals are as follows:

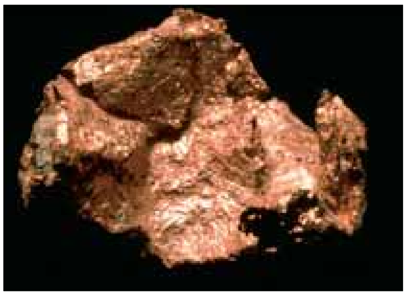
1. They do not have a tendency to have an ionic structure so they are inert in chemical reactions.
2. They have very high density, so they are called heavy metals.
3. They are found in nature as pure metals.

In this chapter, we will focus on some of these inert metals, such as copper, gold, silver, platinum and mercury.

Name	Copper	Silver	Gold	Mercury	Platinum
Symbol	Cu	Ag	Au	Hg	Pt
Atomic number	29	47	79	80	78
Atomic mass	63.5	107.9	197	200.6	195.1
Electron configuration	$[Ar]3d^{10}4s^1$	$[Kr]4d^{10}5s^1$	$[Xe]4f^{14}5d^{10}6s^1$	$[Xe]4f^{14}5d^{10}6s^2$	$[Xe]4f^{14}5d^9 6s^1$
Melting point (°C)	1084.6	961.8	1064.2	-38.8	1768.3
Boiling point (°C)	2927	2162	2856	356.7	3825
Density (g/cm^3)	8.92	10.5	19.3	13.6	21.1
I^{st} Ionization energy (kj/mol)	745.5	731	890.1	1007.1	870
Atomic radius (pm)	138	153	144	149	128
Common oxidation numbers	+1, +2	+1	+1, +3	+1, +2	+2, +4
Color	metallic-reddish	metallic-silver	bright-yellow	silvery white	greyish-white
Physical state at 25°C	solid	solid	solid	liquid	solid
Origin and meaning of name	*cuprum* - Cyprus island	*argentum* - silver	*Aurum* - gold	*hydrargyrum* - liquid silver	*platine* - silver
Earth's crust abundance (%)	6.8×10^{-3}	8.0×10^{-6}	3.1×10^{-7}	6.7×10^{-6}	3.7×10^{-6}

Table 1 *Some properties of Cu, Au, Ag, Pt and Hg.*

1. COPPER



Copper mineral

Copper was the first metal to be used by humans. Today, it's the second most consumed metal, after iron. Copper was found first in Cyprus. That's why it takes its name from the Latin word Cuprum, meaning the island of Cyprus.

The relative atomic mass of copper is 63.54 g/mol, and its electron configuration is $[Ar]3d^{10}4s^1$. Copper is a very soft metal with a characteristic red color. It is the lightest inert metal with a density of 8.92 g/cm^3. It melts at 1084.6°C and boils at 2927°C. After silver, it is the second best conductor of electricity.

The extensive use of copper makes it the second metal in commercial importance, after iron. The chief use of copper is in the production of all types of electrical wiring. Copper is also used in the production of alloys. Some important alloys are:

brass (Cu, 60-82%; Zn, 18-40%), bronze (Cu, 70-95%; Zn, 1-25%; Sn, 1-18%), aluminum bronze (Cu, 90-98%; Al, 2-10%)

1.1 OCCURRENCE

In spite of the wide use of copper, it is not abundant in the earth's crust (1.10^{-4}%). In nature, it is found as compounds and in elemental form.

The most important copper minerals are chalcopyrite (copper pyrite) ($Cu \cdot FeS_2$), chalcocite (Cu_2S), agurite ($CuCO_3$–$Cu(OH)_2$), cuprite (Cu_2O) and malachite ($CuCO_3 \cdot Cu(OH)_2$).

Azurite-malachite

Jasper with quartz and copper

Preparation

In the preparation of copper, chalcopyrite ($CuFeS_2$) is commonly used, since it is the most abundant ore of copper. First the ore is roasted. The Cu_2S formed is converted to liquid form by heating to 1100°C, and the slag is removed.

$$2CuFeS_2(s) + 4O_2(g) \longrightarrow Cu_2S(s) + 3SO_2(g) + 2FeO(s)$$

The Cu_2S is collected and roasted again to remove sulfur as sulfur dioxide.

$$3Cu_2S(s) + 3O_2(g) \longrightarrow 6Cu(s) + 3SO_2(g)$$

1.2 CHEMICAL PROPERTIES

Copper is a less active metal than hydrogen. That's why it doesn't react with water, HCl, diluted H_2SO_4 or other acids. Acids like HNO_3 and concentrated H_2SO_4, which have oxidizing properties, react with copper. In the reactions, H_2 gas is not evolved.

In damp places, copper rusts and the surface is covered by a green colored layer containing $CuCO_3$ and $Cu(OH)_2$. This layer prevents the metal from oxidizing further.

$$2Cu(s) + O_2(g) + CO_2(g) + H_2O(l) \longrightarrow CuCO_3 \cdot Cu(OH)_2(s)$$ (Figure 1).

Copper also reacts with most of the nonmetals when heated.

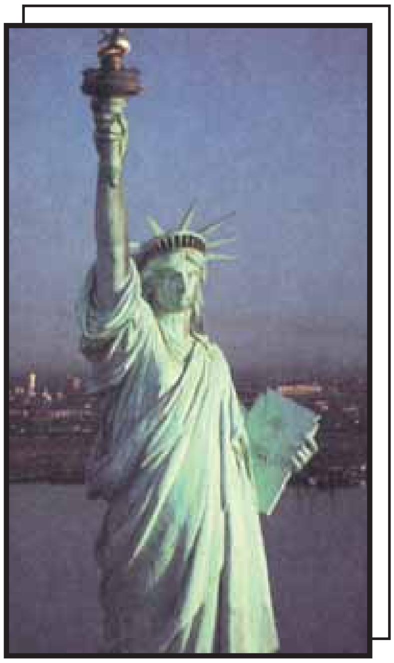


Figure 1 The green color of the statue of Liberty is due to a coating of $CuCO_3$ and $Cu(OH)_2$, produced by the weathering of its copper covering.

Reactions

1. Hot copper reacts with all of the halogens.

$Cu + Cl_2 \longrightarrow CuCl_2$ (at low temperature)

$2Cu + Cl_2 \longrightarrow 2CuCl$ (at high temperature)

2. Oxygen and sulfur produce different oxides and sulfides at different temperatures.

$2Cu + O_2 \longrightarrow 2CuO$ (below 1000°C)

$4Cu + O_2 \longrightarrow 2Cu_2O$ (above 1000°C)

$Cu + S \longrightarrow CuS$ (below 1000°C)

$2Cu + S \longrightarrow Cu_2S$ (above 1000°C)

3. Copper reacts with oxidizing acids in the presence of oxidizing agents such as oxygen and chlorate ion without producing H_2 gas.

$2Cu + 4HNO_3(dil.) + O_2 \longrightarrow 2Cu(NO_3)_2 + 2H_2O$

$2Cu + 2H_2SO_4(dil.) + O_2 \longrightarrow 2CuSO_4 + 2H_2O$

Copper reacts directly with diluted HNO_3 by producing NO gas.

$3Cu + 8HNO_3(dil.) \longrightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$

Reaction of copper with concentrated HNO_3 produces NO_2 gas.

$Cu + 4HNO_3(conc.) \longrightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$ (Figure 2).

Copper with concentrated H_2SO_4 produces SO_2 gas.

$Cu + 2H_2SO_4(conc.) \longrightarrow CuSO_4 + SO_2 + 2H_2O$

Figure 2 Reaction of copper with concentrated HNO_3 produces orange red colored NO_2 gas.

4. Copper displaces less reactive metals from their compounds in aqueous solution.

$Cu + 2AgNO_3 \longrightarrow Cu(NO_3)_2 + 2Ag$ (Figure 3).

No comment.

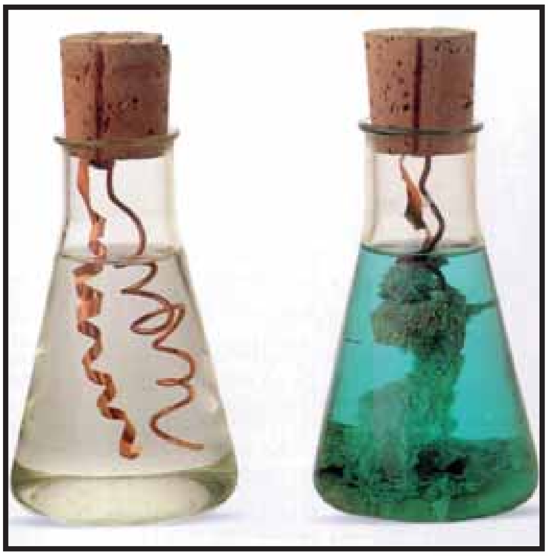


Figure 3 Copper reacts with silver nitrate to give copper nitrate solution and metallic silver.

1.3 COMPOUNDS

Copper has +1 and +2 charges in its compounds. ***Ions with +1 are called copper (I) or cuprous, and ions with +2 are called copper (II) or cupric.***

Cuprous compounds are stable at high temperatures and cupric compounds are stable at low temperatures. When cupric compounds are heated they are transformed into cuprous compounds.

In damp places, copper (I) compounds are oxidized to copper (II) compounds, which contain water and are green and blue colored, while salts of copper (I) are colorless. Solutions of copper salts produce green color in a Bunsen burner. Copper solutions are poisonous.

The most important cuprous compounds are: copper (I) oxide (Cu_2O), and copper (I) chloride (Cu_2Cl_2), and those of cupric compounds are copper (II) chloride ($CuCl_2$), and copper (II) sulfate ($CuSO_4$).

Copper (II) sulfate is found in the crystal hydrate form $CuSO_4 . 5H_2O$. It is the most important compound of copper. If it contains 5 moles of water, it is blue colored. If it looses all water it becomes white. It's obtained by the following reaction.

$$2Cu + O_2 + 2H_2SO_4 \longrightarrow 2CuSO_4 + 2H_2O$$

In industry copper (II) sulfate is used in electrolysis. $CuSO_4$ kills bacteria and microbes. It is used in purifying drinking water and preventing insects and fungi from attacking wood.

Copper (I) species are called as cuprous and copper (II) species are called as cupric compounds.

USES

Most copper is used in the production of copper wire and plates. It's vaporized easily in a vacuum, so it is used in mirrors which reflect infrared rays. It produces strong alloys with gold. In the production of bronze and brass it is used in different ratios. Its soluble compounds are poisonous. Copper (I) oxide a pigment, is used to produce red colored glass.


Some copper tools


Flame test of copper


Copper sulfate has a blue color in its hydrate form. But when it is dried, its color turns white.

2. SILVER

Silver has been used since ancient times. It takes its name from the Latin word argentum, which means silver. Silver is a transition metal with electron configuration $[Kr]4d^{10}5s^1$. It has a characteristic white color. It is a heavy metal with a density of 10.5 g/cm^3. Its melting point is 961.8 °C, and boiling point is 2162 °C. It is the second metal after gold in malleability: a 2 km wire can be obtained from 1 g silver. Although it is the best conductor of heat and electricity of all the metals, silver can not be used for that purpose because of its cost.

Silver is an inert metal so it is used in the jewellery industry. It is also used to plate of metal objects which are more active than silver to prevent their corrosion.

Some silver minerals

2.1 OCCURRENCE

The percentage of silver in nature is very low 8×10^{-6}%. Generally it is found in metallic form and as minerals in small amounts. The most important minerals are: argentite or silver glance (Ag_2S), pyrangyrite or ruby silver (Ag_3SbS_3), silver - arsenic glance (Ag_3AsSi_3) and horne silver (AgCl).

Preparation

Some of the production of silver is from its minerals, but most is from the production of copper and lead as a secondary product. Different methods are used to obtain silver.

a. Amalgam Method

Ores containing free silver are powdered and mercury (Hg) is added. Silver is melted in Hg and, by distilling this mixture, Hg is evaporated. At the end, silver remains in the container and the mercury is condensed and used again.

b. Cyanide Method

This method is used for Ag_2S and some other ores which are low in silver. Powdered ore and molten NaCN are mixed to obtain complex salt of silver and cyanides.

$$Ag_2S + 4NaCN \longrightarrow \underset{\text{sodium silver cyanide}}{2NaAg(CN)_2} + Na_2SO_4$$

By using zinc (Zn), impure silver is obtained. This silver is purified by electrolysis.

$$Zn + 2NaAg(CN)_2 \longrightarrow \underset{\text{sodium zinc cyanide}}{Na_2Zn(CN)_4} + 2Ag$$

2.2 CHEMICAL PROPERTIES

The electron configuration of silver ends with $4d^{10}5s^1$. Thus it has +1 oxidation in its compounds by giving $5s^1$ electron. +2 and +3 oxidation numbers are also possible for some of its compounds.

Silver is an inert metal so it is not affected by the acids, like HCl, which do not have oxidizing property. It reacts with oxidizing acids without producing hydrogen (H_2) gas. In air and also at high temperatures, it is not oxidized. In the presence of H_2S, the surface of silver is covered by black Ag_2S. That's why compounds containing sulfur, like egg and mustard, change the color of silver.

1. Silver in air does not react, but it does react with ozone by forming peroxide.

   $$2Ag + 2O_3 \longrightarrow Ag_2O_2 + 2O_2$$

2. Silver reacts with H_2S in moisture by producing black Ag_2S compound.

   $$4Ag + 2H_2S(aq) + O_2 \longrightarrow 2Ag_2S + 2H_2O$$

3. Silver reacts with oxidizing acid, but with diluted H_2SO_4 it doesn't give any reaction.

$$2Ag + 2H_2SO_4 \text{ (conc.)} \longrightarrow Ag_2SO_4 + SO_2 + 2H_2O$$

$$Ag + 2HNO_3 \text{ (conc.)} \longrightarrow AgNO_3 + NO_2 + H_2O$$

$$3Ag(s) + 4HNO_3(dil.) \longrightarrow 3AgNO_3(aq) + NO(g) + 2H_2O(l)$$

2.3 COMPOUNDS

The most common compound of silver is silver nitrate.

Silver nitrate, $AgNO_3$

It is a colorless and crystalline salt. It is very soluble in water and is obtained by the reaction of metallic silver and nitric acid.

$$3Ag(s) + 4HNO_3(dil.) \longrightarrow 3AgNO_3(aq) + NO(g) + 2H_2O(l)$$

$$Ag(s) + 2HNO_3(conc.) \longrightarrow AgNO_3(aq) + NO_2(g) + H_2O(l)$$

$AgNO_3$ is decomposed when heated.

$$2AgNO_3(s) \xrightarrow{\text{heat}} 2Ag(s) + 2NO_2(g) + O_2(g)$$

When $AgNO_3$ reacts with organic substances like skin, black colored metallic silver is formed.

USES

Silver is widely used in the chemical industry despite not mixing easily with other substances. Powdered silver is an excellent catalyst used in industrial alcohol production and in production of fertilizers with ammonia compounds. Silver iodine compounds have been used to "seed" clouds, encouraging them to produce rain. This is not only useful in wiping out drought but also in taking the violence out of hurricanes before they reach populated areas. In the right conditions, 28 grams of silver iodide could produce 12 million liters of rain.

Silver jewelry

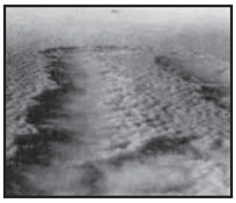
Silver is useful in medicine for absorbing the oxygen that germs need to live, there by killing them. Surgeons' instruments are often made from silver to reduce the chance of infection, and silver wire and plates can be used to repair broken limbs. Tooth fillings are made out of a silver, tin and mercury alloy. Silver can also be used to purify water absorbing up to 20 times its own weight in water and killing the germs. This provides astronauts their supply of fresh water, and it can also be used by campers who get their water from streams and rivers.

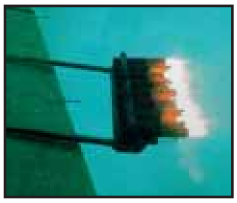
Ag_2O is used in mirror preparation, AgBr in photography, and $AgNO_3$ is used in medicines for eyes.

EXTRA

Modification of The Weather By Cloud Seeding

Ever since man began to understand and investigate the weather, he has tried to find ways to control it. From witch doctors and rain dances, we have come a long way to now trying to influence the timing of rain and its location by modifying the weather. Man is now investigating increasing precipitation, dispersing fog and weakening hurricanes by cloud seeding.





Seeding of tropical cumulus clouds, and indeed any clouds, requires that they contain supercooled water. That is, liquid water is colder than zero Celsius.

Introduction of a substance, such as silver iodide, that has a crystalline structure similar to that of ice will induce freezing. In mid-latitude clouds, the usual seeding strategy has been based upon the vapor pressure being lower over water than over ice. When ice particles form in supercooled clouds, they grow at the expense of liquid droplets and become heavy enough to fall as rain from clouds that otherwise would produce none.

Photography

40% of all silver produced is used in photography. In this industry silver halides are used in an emulsion with gelatine to make the film. When the film is exposed, the halide ion absorbs light and releases an electron. This electron is taken by silver ion to give silver metal. These metal particles produce the image.

$$2I^- \rightarrow I_2 + 2e^- \qquad 2Ag^+ + 2e^- \rightarrow 2Ag$$

3. GOLD

Gold is a soft yellow metal, while other metals (except copper) are silvery in appearance. It is the most malleable and ductile of all metals. Gold foil can be prepared by hammering the metal into very thin sheets. 1 gram of gold can be drawn into a wire more than 3 km in length. Pure gold is too soft to be used for jewelry and coins. For such purposes it is always alloyed with copper, silver or some other metal. The purity of gold is expressed in karats, a designation that indicates the number of parts, by weight, of gold in 24 parts of alloy. Thus 24– karat gold is the pure metal, while a 10– karat alloy is 10/24 gold by weight. Red or yellow gold alloys contain copper, and white gold contains palladium, nickel or zinc.

EXTRA

What Is 24 Karat Gold?

The term "karat" or carat refers to the percentage of gold versus the percentage of an alloy in an object. Gold is too soft to be usable in its purest form. It has to be mixed with other/metals.

Karats	% fine gold
24	100
22	91.75
18	75
14	58.5

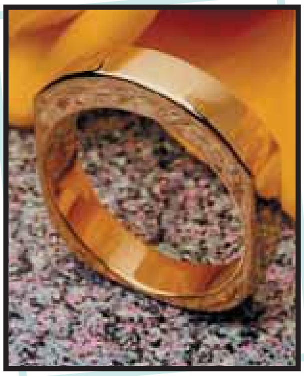
What is aqua regia?

"Aqua regia," also known as nitrohydrochloric acid, is a mixture of one part concentrated nitric acid and three parts concentrated hydrochloric acid. The chemical reaction between the acids makes it possible to dissolve all metals except silver. The reaction of metals with nitrohydrochloric acid typically involves oxidation of the metals to nitric oxide. The term comes from Latin and means royal water.

3.1 OCCURRENCE

In the earth's crust, gold is found in 3.1 x 10^{-7}%. In nature, it is found in the elemental form. The most important minerals of gold are calavarite ($AuTe_2$) and silvanite ($AuAgTe_4$).

When people discovered the convenience of carrying small coins with them to trade rather than goods to barter, gold was chosen because small pieces could carry relatively high values. This process was disrupted by the first world war, and gold was used less and less as actual currency, overtaken by paper currencies.



A gold ring

Preparation

In nature, gold is found in the elemental form with some impurities. To obtain pure gold, first impure gold is dissolved in aqua regia. When the solution obtained is evaporated, yellow crystals of chloroauric acid ($HAuCl_4 \cdot 4H_2O$) are formed. When this coordination compound is heated, hydrogen chloride is evolved, and red crystalline gold(III) chloride ($AuCl_3$) remains. When this compound is heated to 175°C, it decomposes to gold (I) chloride (AuCl) and at higher temperatures the metal is obtained.

$$HAuCl_4 \xrightarrow{\text{heat}} AuCl_3 + HCl$$

$$AuCl_3 \xrightarrow{\text{heat}} AuCl + Cl_2$$

$$2AuCl \xrightarrow{\text{heat}} 2Au + Cl_2$$

Sometimes gold is obtained from its ores. The proper method of preparation of gold from ores is the cyanide method mentioned before for silver.

3.2 CHEMICAL PROPERTIES

Gold is a very inactive metal. It neither combines directly with oxygen nor corrodes in the atmosphere. The metal is not affected by any single common acid or by bases. However, it dissolves readily in aqua regia.

Now we will focus on some important reactions of gold.

Reactions

1. Reaction of gold with aqua regia (king water)

   $Au + HNO_3 + 4HCl \longrightarrow HAuCl_4 + NO\uparrow + 2H_2O$

2. At high temperature, it reacts with fluorine, chlorine and bromine, forming water soluble gold (III) halides.

   $2Au + 3X_2 \longrightarrow 2AuX_3$ (X = F, Cl and Br)

   When these compounds are hydrolyzed in water, they produce hydrates of gold (III) oxide, $Au_2O_3 \cdot 4H_2O$, which have a brown color.

3.3 COMPOUNDS

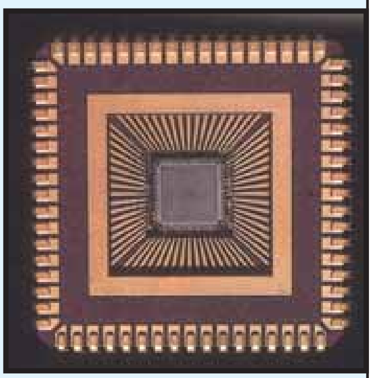
In most of its compounds, gold has an oxidation number of +1 or +3. All compounds of gold are decomposed by heating, because they are thermally unstable. Gold(I) halides undergo auto-oxidation-reduction reactions in water, forming gold(III) halides and the metal. Gold forms two oxides, Au_2O and Au_2O_3, and the corresponding hydroxide, AuOH (a weak base). Potassium cyanide reacts with gold (I) and gold (III) compounds, giving the complex soluble salts $Na[Au(CN)_2]$ and $Na[Au(CN)_4]$, which are important in the extraction of gold from its ores and in gold plating operations.

USES

Gold is used in microchips

Because of the chemical stability of gold, it is used in technology, to coat metals to prevent their corrosion, and fill teeth. As a good conductor, it is used in transistors and in connections of electrical circuits.

Every year tons of gold are used in the production of jewellery. Gold is a valuable element, and is used as a money transfer between banks and treasuries.



Gold is used in microchips

4. MERCURY

Mercury is the last element of group 2B and ends with the electron configuration $5d^{10}6s^2$. It is the only metal which is liquid at room temperature. In Latin, "mercury" means liquid silver. Its melting point is –38.8 °C and boiling point is 356.7°C. It is a silvery white colored metal with density of 13.6 g/cm^3. It has a very high boiling point, so it has a very low vapor pressure at room temperature with respect to other liquids. Its vapor is very toxic. Like all the other metals, in liquid or solid form it conducts electricity.

The most important property of mercury is to combine with all metals to by form alloys, except iron and platinum. These alloys are called amalgams.

Examples are silver amalgam (Ag-Hg) and cadmium amalgam (Cd-Hg). Amalgams may be liquid or solid. Generally, Hg rich amalgams are liquid.

Pure mercury

The alchemists were generally interested in mercury because it is a heavy and silvery metal like silver. But it is liquid at room temperature. They thought that if they could only harden it, it would be real silver. So it was called "quicksilver" or "live silver".

We must be very careful if mercury filled thermometers or other instruments containing mercury are broken, because mercury is volatile and poisonous to living organisms.

4.1 OCCURRENCE

Mercury is found on the surface of rocks in elemental form and at the same time as the alloys of silver and gold. The most abundant type of mercury mineral in nature is HgS mercury (II) sulfide, called cinnabar.

Preparation

In the preparation of mercury, its most common ore, cinnabar, is used. When the ore is roasted in air, the sulfide oxidizes to SO_2 and mercury is set free as a vapor. The vapor is then condensed by distillation.

$$HgS + O_2 \xrightarrow{\text{heat}} Hg + SO_2$$

Mercury is also obtained by heating HgS in the presence of calcium oxide. The Hg produced is not pure.

$$4HgS + 4CaO \longrightarrow 4Hg + 3CaS + CaSO_4$$

4.2 CHEMICAL PROPERTIES

Mercury has the electron configuration $[_{54}Xe]4f^{14}5d^{10}6s^2$. In its compounds it has +1 and +2 oxidation numbers. In reactions with acids it does not produce hydrogen gas. Although it is stable in air, it forms its oxides in trace amount.

Reactions

Mercury is not affected by water, water vapor, HCl and diluted H_2SO_4. However, it reacts with concentrated H_2SO_4 to produce SO_2 gas.

$$Hg + 2H_2SO_4(conc.) \longrightarrow HgSO_4 + SO_2\uparrow + 2H_2O$$

Mercury reacts with concentrated nitric acid, producing mercury (II) nitrate and with diluted HNO_3, producing mercury (I) nitrate.

$$Hg + 4HNO_3(conc.) \longrightarrow Hg(NO_3)_2 + 2NO_2\uparrow + 2H_2O$$

$$3Hg + 4HNO_3(dil) \longrightarrow 3HgNO_3 + NO\uparrow + 2H_2O$$

Hg^{+} compounds are called mercurous compounds. Hg^{2+} compounds are called mercuric compounds.

$Hg_2(NO_3)_2$ mercurous nitrate, $Hg(NO_3)_2$ mercuric nitrate

4.3 COMPOUNDS

1. Oxides

Mercury (II) oxide, HgO is a yellow-orange powder. It is formed by the reaction of mercury (I) salts with strong base solutions.

$$Hg_2(NO_3)_2(s) + 2NaOH(aq) \longrightarrow HgO(s) + Hg(l) + 2NaNO_3(aq) + H_2O(l)$$

Mercury (II) oxide is obtained by two different means.

The first one is the heating of metallic Hg in air at 350°C. The product of the reaction is yellow-orange colored HgO.

$$2Hg + O_2 \xrightarrow{350°C} 2HgO$$

The second way is to add a strong base to mercury (II) salts.

$$HgCl_2 + 2NaOH \longrightarrow HgO + 2NaCl + H_2O$$

The HgO obtained decomposes on heating

$$HgO(s) \xrightarrow{heat} Hg(l) + 1/2O_2(g)$$

2. Halides

Mercury (II) halides can be obtained by the reaction of mercury (II) salt solutions with sodium halides.

$$Hg(NO_3)_2 + 2NaX \longrightarrow HgX_2 + 2NaNO_3$$

The most important halides of mercury are $HgCl_2$ and Hg_2Cl_2.

Hg_2Cl_2 is obtained by direct reaction of Hg with $HgCl_2$.

$$HgCl_2 + Hg \longrightarrow Hg_2Cl_2$$

In addition to these compounds mercury (II) nitrate ($Hg(NO_3)_2$) and mercury (II) sulfide (HgS) are also important compounds of mercury.

USES

Mercury is used in many different areas. It has a low melting point and a high boiling point, so it is used in thermometers and barometers. Mercury is also used in the production of batteries for electronic instruments. The most important use of Hg is in dentistry, for filling teeth. The alloys of mercury (amalgams) with silver, lead, cadmium and copper are used for that purpose.

Mercury is a very dangerous metal for living organisms. Because of this mercury-free batteries are getting more popular every day.

A crop duster spraying pesticides on corn.

Thermometer

5. PLATINUM

Spanish scientist Antonio de Wloa examined the gold ores brought from Colombia in 1735, and found a metal with similar properties to gold. He called this metal platina, from the spanish word that means platinum. Platinum has the atomic number 78 and electron configuration $[Xe]4f^{15}d^{9}6s^{1}$. It has +2 and +4 oxidation numbers in its compounds.

Platinum is white, malleable and very ductile. It has a density of $21.2 g/cm^3$. Its melting point is 1768.3°C and boiling point is 3825°C.

5.1 OCCURRENCE

Generally platinum is found in elemental form, but it can also be found as compounds like Pt As_2 (sperilite). It is also found in copper - nickel salts.

It is obtained in laboratories by heating ammonium salts.

$$(NH_4)_2PtCl_6 \xrightarrow{300°C} Pt + 2NH_4Cl + 2Cl_2\uparrow$$

5.2 CHEMICAL PROPERTIES

Platinum does not react with acids, bases and water. It becomes more active when heated, and reacts with some nonmetals.

Reactions

It is very difficult to oxidize platinum but when it is oxidized, platinum (II) oxide PtO is formed.

$$2Pt + O_2 \longrightarrow 2PtO$$

It can react only with king water, a mixture of HCl and HNO_3, giving H_2PtCl_6 hexachloro platinate acid.

$$3Pt + 18HCl(conc.) + 4HNO_3(conc) \longrightarrow 3H_2PtCl_6 + 4NO\uparrow + 8H_2O$$

Platinum reacts with fluorine and chlorine by heating. As a result of these reactions, platinum (IV) fluoride (PtF_4) and platinum (IV) chloride ($PtCl_4$) are obtained, respectively.

$$Pt + 2F_2 \xrightarrow{450°C} PtF_4$$

$$Pt + 2Cl_2 \xrightarrow{300°C} PtCl_4$$

5.3 COMPOUNDS

Important compounds of platinum are platinum (IV) fluoride (PtF_4), platinum (IV) chloride ($PtCl_4$) and hexachloro platinate acid, H_2PtCl_6.

PtF_4 has a yellowish brown color and $PtCl_4$ has a reddish brown color.

H_2PtCl_6 is a strong acid. Its solution has a reddish brown color.

USES

Platinum is used as a catalyst in most organic reactions because it absorbs hydrogen easily. Platinum is also used in the oxidation of NH_3 to NO, and in the preparation of very thin plates. It's very inactive, so it's used in filling teeth, in the production of medical instruments, and in jewelry. In industry, powdered platinum is used as a catalyst in many reactions.

Platinum resistance wires are used for constructing high-temperature electric furnaces.

In the military, platinum is used for coating missile nose cones and jet engine fuel nozzles.

SUPPLEMENTARY QUESTIONS

1. Compare the electrical conductivity of Cu and Fe.

2. What is the color of copper metal after corrosion?

3. What are the oxidation numbers of copper in its compounds.

4. Complete the following reactions.

a. $Cu_2S(s) + O_2(g) \longrightarrow$ ....

b. $Cu + Cl_2 \longrightarrow$ .... (at high T)

c. $Cu + Cl_2 \longrightarrow$ .... (at low T)

d. $Cu + HNO_3(dilute) \longrightarrow$ ....

e. $Cu + HNO_3(conc.) \longrightarrow$ ....

7. What color do the solutions of copper salts give in flame test?

8. Silver is used in many different places in our daily life. Where do we use the following silver compounds?

Ag_2O : ....

AgBr : ....

$AgNO_3$: ....

9. Write two preparation methods of silver.

10. Generally which oxidation state does silver have in its compounds?

11. Which compound is formed on the surface of silver metal in the presence of H_2S?

12. Complete the following reactions.

a. $Ag + O_3 \longrightarrow$

b. $Ag + H_2SO_4(conc.) \longrightarrow$

c. $Ag + HNO_3(diluted) \longrightarrow$

d. $Au + HNO_3 + 3HCl \longrightarrow$

e. $HgS + O_2 \longrightarrow$

13. What are the most important ores of gold?

14. Which oxidation states does gold have in its compounds?

15. Research gold-copper alloys and find the different names of these alloys.

16. What is the most common oxidation state of platinum?

17. What is the common name of mercury alloys?

18. What is the most common oxidation state of mercury?

19. What are the chief ores of copper.

20. **a.** How is refined copper prepared?

b. Described chemical characteristics of copper with chemical reactions involved.

21. Research which noble metal is more expensive than the others?

22. Try to find meaning of words such as ounce, carat.

23. Compare the color of noble metals with color of other metals.

MULTIPLE CHOICE QUESTIONS

1. Which one of the following is not an inert metal?

A) Cu B) Ag C) Au D) Pt E) Mn

2. Which one of the following metal is in liquid state at room temperature?

A) H_2 B) He C) Ag D) Hg E) Au

3. Which one of the following takes only one kind of charge?

A) Ag B) Au C) Cu D) Pt E) Hg

4. Which one of the following was discovered in Cyprus?

A) Silver B) Copper C) Gold D) Platinum E) Mercury

5. Which one of the following has the lowest melting point?

A) Ag B) Au C) Hg D) Pt E) Cu

6. What are the components of brass?

A) Cu-Pb B) Al-Zn C) Al-Cu D) Cu-Zn E) Fe-Pb

7. What are the components of bronze?

A) Cu-Pb B) Cu-Sn C) Fe-Al D) Zn-Cu E) Fe-Pb

8. Which one of the following acid reacts with inert metals?

A) H_3BO_3 B) HI C) HF D) HCl E) H_2SO_4

9. Which one of the following gases is produced at the end of the reaction?

$3Cu(s) + 8HNO_3(dil) \longrightarrow$

A) NO B) N_2O C) NO_2 D) H_2 E) OH

10. Ag conducts electricity better than Cu. Why is Cu preferred instead of Ag?

A) Due to finance B) Due to hazards

C) Due to color D) Due to more reaction

E) Due to conductivity of heat

11. Which one of the following compounds gives the red color to the glass?

A) CuO B) CuCl C) Cu_2O D) $CuSO_4$ E) $CuCO_3$

12. Which one of the following compound is used on photography?

A) $AgNO_3$ B) AgI C) Ag_2O D) AgBr E) AgCl

13. What is karat of pure gold?

A) 12 B) 14 C) 18 D) 22 E) 24

14. Which one of the following is formula of king water?

A) $3HNO_3 + HCl$ B) $HNO_3 + 3HCl$ C) $HNO_3 + HCl$

D) $H_2SO_4 + 3HCl$ E) $H_2SO_4 + HCl$

15. Which one is the most abundant metal in the earth's crust?

A) Cu B) Ag C) Au D) Pb E) Hg

16. Which one of the following compounds is used in cloud seeding?

A) Ag_2O B) $AgNO_3$ C) AgI D) AgCl E) Ag_2SO_4

17. Which one of the following is an amphoteric compound of gold?

A) $AuCl_3$ B) AuF_3 C) $AuTe_2$ D) $Au(OH)_3$ E) Au_2O

18. What is the oxidation state of platinum in its compounds?

A) +1 B) +3 C) –3 D) –4 E) +4

19. Which one of the following metals is the basic substance to form amalgam?

A) Hg B) Pb C) Al D) Ag E) Au

20. Which one of the following is the most toxic?

A) Ag B) Hg C) Pb D) Cu E) Au

21. Which one of the following uses of copper is most closely related to the fact that copper is resistant to corrosion?

A) The construction of hot water tanks

B) The manufacture of cooking pots

C) The production of electrical wire

D) The preparation of various alloys

E) The making of jewellery

PUZZLE

These words have been misspelt, but are anagrams. Luckily, clues have been left and these are in the proper order. Put the correct spellings in the spaces provided.

1.	EAMTLS	These elements are on the left side of the periodic table.	____________
2.	IRNCIT	This acid is made industrially by the Ostwald process.	____________
3.	RIULSHUPC	This acid is made industrially by the Contact process.	____________
4.	DNIOAIDT	A reaction in which two molecules react together to form a single larger molecule.	____________
5.	SOSL	They production of metal ions involves the _______ of electrons from the metal involving.	____________
6.	EXDRO	This describes a chemical reaction involving oxidation and reduction.	____________
7.	LUPDESHI	A compound made of a metal and sulphur.	____________
8.	XEOGNY	A reactive oxidising agent making up 21% of the atmosphere.	____________
9.	CONHIERL	A poisonous, choking, oxidising agent which is a member of the halogen group.	____________
10.	IRMDHOECAT	When this potassium salt acts as an oxidising agent its anion is reduced from $Cr_2O_7^{2-}$ to Cr^{3+}	____________
11.	NAEGID	Electrons are ________ by an atom when it is converted into an anion.	____________
12.	EUCRDIOTN	A chemical reaction in which a compound loses oxygen.	____________
13.	BCRAON	A reducing agent, which can reduce any metal.	____________
14.	HPULSRU DIOIDEX	This gas forms sulphurous acid when dissolved in water	____________
15.	XIPOEEDR	This type of compound contains oxygen, but with an oxidation number of –1, not –2, as oxygen usually has.	____________
16.	ONCTNTERAEDC	The strength of sulphuric or nitric acid needed to react with carbon.	____________
17.	ODSNXGIII	This agent accepts electrons from another substance.	____________
18.	XIADOTINO	A chemical reaction in which a compound loses hydrogen.	____________
19.	CETOERNLS	When these are lost the atom is oxidised.	____________
20.	RISVLE	The best conductor of electricity.	

ANSWERS

SUPPLEMENTARY QUESTIONS

ALKALI METALS

ALKALINE EARTH METALS

7. A : 5.25 g
8. A : 4.8 g
9. A : 8.85 g
10. A : 12.8 g
11. A : 1.12 L
12. A : 22.4 L

SOME TRANSITION METALS

15. Fe_3O_4
16. 18.2 g
17. 38 g
18. 10.7 g

ALUMINUM

1. 33.6 L
3. 0.4 mol
4. 5.6 mol
5. 10.26 g
14. 1

LEAD

INERT METALS

MULTIPLE CHOICE

ALKALI METALS

1. B	5. A	9. C	13. A	17. C	21. A
2. A	6. E	10. C	14. D	18. B	
3. E	7. D	11. E	15. D	19. E	
4. C	8. C	12. B	16. A	20. C	

ALKALINE EARTH METALS

1. B	6. A	11. A	16. C	21. E	26. D
2. C	7. B	12. C	17. B	22. C	27. A
3. A	8. E	13. C	18. D	23. C	28. D
4. E	9. B	14. C	19. C	24. D	
5. D	10. A	15. C	20. E	25. A	

SOME TRANSITION METALS

1. A	5. B	9. C	13. C	17. E
2. E	6. E	10. E	14. D	18. B
3. C	7. A	11. B	15. A	19. C
4. D	8. D	12. A	16. E	20. D

ALUMINUM

1. B	5. D	9. A	13. B	17. A
2. B	6. E	10. E	14. C	18. A
3. C	7. E	11. A	15. A	19. D
4. D	8. B	12. D	16. D	20. C

LEAD

1. B	3. B	5. C	7. C	9. D	11. E
2. D	4. B	6. C	8. C	10. D	

INERT METALS

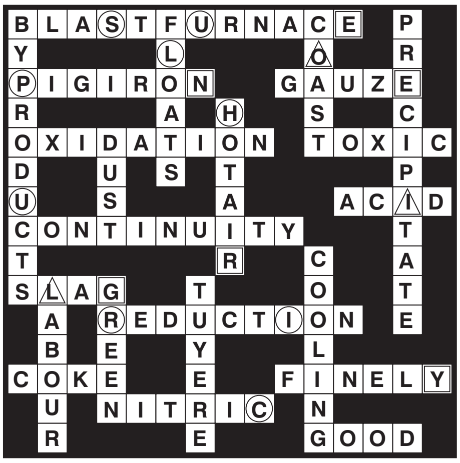
1. E	5. C	9. A	13. E	17. D	21. A
2. D	6. D	10. A	14. B	18. E	
3. A	7. B	11. C	15. A	19. A	
4. B	8. E	12. D	16. C	20. B	

PUZZLE

ALKALI METALS

Words	ACIDIC	ONE
DISPLACEMENT	UNIVERSAL	LITHIUM
CESIUM	CLASSIFICATION	DÖBEREINER
SHINY	GASES	CATALYSTS
PERIODS	INFLAMMABLE	
AIR	VULCANISE	*Scientists*
POTASSIUM	COMBINATION	MEYER
COLOURED	CONDUCT	MENDELEEV
REACTIONS	TRANSITION	MOSELEY

TRANSITION METALS



Extra Clues

◯ **SULPHURIC**

☐ **ENERGY**

△ **OIL**

ALKALINE EARTH METALS

A	B	C	D	E	F	G	H	I	J	K	L	M	N	O
5	10	16	19	1	8	21	11	2	17	7	18	28	6	23

P	Q	R	S	T	U	V	W	X	Y	Z	AA	BB	CC	DD
13	25	3	27	9	22	4	30	15	26	12	29	24	14	20

INERT METALS

1. METALS
2. NITRIC
3. SULPHURIC
4. ADDITION
5. LOSS
6. REDOX
7. SULPHIDE
8. OXYGEN
9. CHLORINE
10. DICHROMATE
11. GAINED
12. REDUCTION
13. CARBON
14. SULPHUR DIOXIDE
15. PEROXIDE
16. CONCENTRATED
17. OXIDISING
18. OXIDATIONS
19. ELECTRONS
20. SILVER

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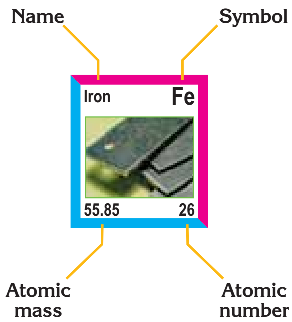
PERIODIC



1 (1A)	2 (2A)		3	4	5	6	7	8	9
Hydrogen H 1.01 1									
Lithium Li 6.94 3	Beryllium Be 9.01 4								
Sodium Na 22.99 11	Magnesium Mg 24.31 1								
Potassium K 39.10 19	Calcium Ca 40.08 20		Scandium Sc 44.96 21	Titanium Ti 47.88 22	Vanadium V 50.94 23	Chromium Cr 52.00 24	Manganese Mn 54.94 25	Iron Fe 55.85 26	Cobalt Co 58.93 27
Rubidium Rb 85.47 37	Strontium Sr 87.62 38		Yttrium Y 88.91 39	Zirconium Zr 91.22 40	Niobium Nb 92.91 41	Molybdenum Mo 95.94 42	Technetium Tc (98) 43	Ruthenium Ru 101.07 44	Rhodium Rh 102.91 45
Cesium Cs 132.91 55	Barium Ba 137.33 56	Lanthanides * 57-70	Lutetium Lu 174.97 71	Hafnium Hf 178.49 72	Tantalum Ta 180.95 73	Tungsten W 183.84 74	Rhenium Re 186.21 75	Osmium Os 190.2 76	Iridium Ir 192.22 77
Fancium Fr {223} 87	Radium Ra (226) 88	Actinides ** 89-102	Lawrencium Lr (262) 103	Rutherfordium Rf (261) 104	Dubnium Db (262) 105	Seaborgium Sg (266) 106	Bohrium Bh (264) 107	Hassium Hs (269) 108	Meitnerium Mt (268) 109

* Lanthanides	Lanthanum La 138.91 57	Cerium Ce 140.12 58	Praseodymium Pr 140.91 59	Neodymium Nd 144.24 60	Promethium Pm (145) 61	Samarium Sm 150.36 62
** Actinides	Actinium Ac (227.03) 89	Thorium Th 232.04 90	Protactinium Pa 231.04 91	Uranium U 238.03 92	Neptunium Np (237.05) 93	Plutonium Pu (244) 94

TABLE

10	11	12	13 (3A)	14 (4A)	15 (5A)	16 (6A)	17 (7A)	18 (8A)
								Helium He 4.0026 2
			Boron B 10.81 5	Carbon C 12.01 6	Nitrogen N 14.0067 7	Oxygen O 15.9994 8	Fluorine F 18.9984 9	Neon Ne 20.18 10
			Aluminum Al 26.98 13	Silicon Si 28.09 14	Phosphorus P 30.97 15	Sulfur S 32.07 16	Chlorine Cl 35.45 17	Argon Ar 39.95 18
Nickel Ni 58.69 28	Copper Cu 63.55 29	Zinc Zn 65.41 30	Gallium Ga 69.723 31	Germanium Ge 72.64 32	Arsenic As 74.92 33	Selenium Se 78.96 34	Bromine Br 79.90 35	Krypton Kr 83.80 36
Palladium Pd 106.42 46	Silver Ag 107.87 47	Cadmium Cd 112.41 48	Indium In 114.817 49	Tin Sn 118.71 50	Antimony Sb 121.76 51	Tellurium Te 127.60 52	Iodine I 126.90 53	Xenon Xe 131.29 54
Platinum Pt 195.08 78	Gold Au 196.97 79	Mercury Hg 200.59 80	Thallium Tl 204.383 81	Lead Pb 207.19 82	Bismuth Bi 208.980 83	Polonium Po (209) 84	Astatine At (210) 85	Radon Rn (222) 86
Darmstadium Uun (271) 110	Unununium Uuu (272) 111	Ununbium Uub (285) 112	113	Ununquadium Uuq 289 114	115	Ununhexium Uuh 292 116	117	118

Europium Eu 151.96 63	Gadolinium Gd 157.25 64	Terbium Tb 158.92 65	Dysprosium Dy 162.50 66	Holmium Ho 164.93 67	Erbium Er 167.26 68	Thulium Tm 168.93 69	Ytterbium Yb 173.04 70
Americium Am (243) 95	Curium Cm (247) 96	Berkelium Bk (247) 97	Californium Cf (251) 98	Einsteinium Es (252) 99	Fermium Fm (257) 100	Mendelevium Md (258) 101	Nobelium No (259) 102