AAAJI OKOMI SILAS

CHM 104 — Inorganic Chemistry

Chemistry of Selected metals and non-metals

$CaSiO_3$ slag

Introduction:

The periodic table can be classified into metals, non-metals and metalloids. Today, about 118 chemical elements are known. However, only 103 of them are well characterized in terms of their properties. The systematic classification of these 103 elements reveals that 79 of them are metals, 17 are non-metals and 7 are metalloids. Metals differ from non-metals in many respects. In fact, metals and non-metals are two extremes as regards their properties. Some of their general properties are outlined below;

General properties of metals and non-metals

Metals	Non-metals
Physical properties	
1) Metals are malleable and ductile	Non-metals are not malleable, not ductile.
2) Metals are strong, not brittle and have good lustre	Non-metals are not strong, generally brittle, poor lustre.
3) Metals are good conductors of electricity	Non-metals are bad conductors of electricity
4) Metals have high tensile strength	Non-metals have poor tensile strength.

(1)

Overall rust formation ~~$Fe_2O_3 + xH_2O \rightarrow 2Fe_2O_3$~~ xH_2O

$$4Fe + 3O_2 + xH_2O \longrightarrow 2Fe_2O_3\, xH_2O$$

$Fe_2SO_4\, Al_2(SO_4)_3\, 24H_2O$

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5) Metals have high melting and boiling points | Non-metals have low melting and boiling points.

6) Other than mercury, all metals are solids at room temperature. | Non-metals can be solids, liquids or gaseous at room temperature.

Chemical properties

1) Metals form electropositive ions in reactions. | Non-metals form electronegative ions in reactions.

2) Form oxides that are basic nature | Form oxides that are either acidic or neutral.

3) Most metals react with dilute acids to give a metal salt and hydrogen gas. | Non-metals do not react with dilute acids.

4) Most metals react with water to give metal hydroxides and hydrogen gas | Non-metals do not react with water.

5) Metals are reducing agents as they donate electrons. | Non-metals are oxidizing agents as they accept electrons.

The metals selected for this study are Sodium (Na), Iron (Fe) and Aluminium (Al). The non-metals are; Oxygen (O) and Nitrogen (N).

Sodium (Na)

Sodium is a soft, silvery, highly reactive element, a member of the alkali metals within (group 1). It has only one stable istope, ^{23}Na. Owing to its reactivity Na is never found pure in nature. it quickly oxidize in air so it must be stored in an inert environment such as kerosene or mineral oil.

Continuation

Occurrence and Extraction

Sodium occurs as sodium chloride (NaCl) in huge quantities in underground deposits and sea water. It is manufactured by electrolysis of fused sodium chloride mixed with fused $CaCl_2$ using the Down's electrolytic cell (Down's process). $CaCl_2$ is added in order for the electrolytic process to be carried out at a much lower temperature (600°c) than 781°c if pure NaCl is used. The Down's electrolytic cell has a graphite anode and a circular iron cathode. with proper outlets. Chlorine is liberated at the anode while sodium metal is deposited at the cathode.

Electrolyte ionizes thus: $NaCl \rightarrow Na^+ + Cl^-$

At the cathode $Na^+ + e^- \rightarrow Na$

At the anode $2Cl^- \rightarrow Cl_2 + 2e$

Bio 102	15
CHM 102	10
CHM 104	10
PHY 102	15
MTH	15
GST 102	10
GST 104	10
Bio 108	5
CHM 108	5
PHY 108	5

Chemical Reactivity of Na

1) Reaction with air (Oxygen): When heated in air, Na burns violently to form the monoxide and the Peroxide

$$4Na_{(s)} + O_{2(g)} \longrightarrow 2Na_2O_{(s)}$$

$$2Na_{(s)} + O_{2(g)} \longrightarrow Na_2O_{2(s)}$$

2) Reaction with water: Na decomposes water vigorously, liberating hydrogen and forming sodium hydroxide.

$$2Na_{(s)} + 2H_2O_{(l)} \longrightarrow 2NaOH_{(l)} + H_{2(g)}$$

The reaction is exothermic

3) Reaction with ammonia: Na reacts with ammonia to give hydrogen gas and a white solid called sodamide.

$$2Na_{(s)} + 2NH_{3(g)} \longrightarrow \underset{\text{(sodamide)}}{2NaNH_{2(s)}} + H_{2(g)}$$

4) Reaction with acids: Na displaces hydrogen from dilute acids and produce the corresponding salts

$$2HCl_{(l)} + 2Na_{(s)} \longrightarrow 2NaCl_{(aq)} + H_{2(g)}$$

5) Reducing action: Na Reduces many compounds when heated with them in the absence of air.

$$Al_2O_3 + 6Na \longrightarrow 2Al + 3Na_2O$$

$$SiO_2 + 4Na \longrightarrow Si + 2Na_2O$$

6) Reactions with the halogens: Na reacts vigorously with all the halogens to form sodium halides.

$$2Na_{(s)} + F_{2(g)} \longrightarrow 2NaF_{(s)}$$

$$2Na_{(s)} + Cl_{2(g)} \longrightarrow 2NaCl_{(s)}$$

~~Na react with H_2O (oxygen)~~

~~$2Na + 2H_2O \longrightarrow 2NaOH + H_2$~~

~~Na react with water to give;~~

~~$Na + H_2$~~

Compounds of Sodium

Some important Compounds of sodium are;
$NaOH$, Na_2CO_3, $NaCl$, $NaHCO_3$.

1) Sodium hydroxide ($NaOH$) is produced Industrially by electrolysis of brine (aqueous solution of $NaCl$).
⇒ In the Lab, $NaOH$ can be prepared by two methods;

i) Action of slaked lime (Calcium hydroxide) on sodium trioxocarbonate (IV) (washing soda).

$$Na_2CO_3(aq) + Ca(OH)_2(aq) \longrightarrow CaCO_3 + 2NaOH$$

ii) Action of Sodium on water.

$$2Na + 2H_2O \longrightarrow 2NaOH + H_2$$

2) Sodium trioxocarbonate (IV) Na_2CO_3; used in soap making, softening of hard water and in manufacture of glass. It is prepared in the laboratory by passing excess CO_2 into $NaOH$;

i) $NaOH + CO_2 \longrightarrow NaHCO_3$.

ii) $NaHCO_3$ is then heated to give the anhydrous Na_2CO_3.

$$2NaHCO_3 \longrightarrow Na_2CO_3 + H_2O + CO_2.$$

3) Sodium hydrogen trioxocarbonate iv; used in fire extinguishers, as baking powder, and in health salts and fizzy drinks. It is prepared in the lab as described in 2(i) above.

IRON (Fe)

Iron is a metal in the first transition series, the fourth most abundant of all the elements. It exists in a wide range of oxidation states -2 to +6, but the +2 and +3 are the most common. Naturally occuring Iron consists of four stable isotopes: ^{54}Fe, ^{56}Fe, ^{57}Fe, ^{58}Fe. The most abundant iron isotope is ^{56}Fe with an abundance of 91.75%.

Occurence and Extraction: Iron occurs chiefly as; haematite (Fe_2O_3), Magnetite (lodestone) (Fe_3O_4) and as Iron pyrites (FeS_2).

The Process of extraction requires first smelting (to obtain the crude metal) and then refining. Iron ore (usually Haematite) is mixed with coke and limestone and heated in blast furnace, at a temperature of about 800°C; Calcium oxide, formed by thermal decomposition of limestone, reacts with the silicon oxide present in Sand, a major impurity in iron ores; to form slag. CO is then used to reduce the oxide to molten iron.

Overall, the processes can be summarized as follows:

- decomposition of limestone

$$CaCO_3 \longrightarrow CaO + CO_2$$

- Formation of slag; CaO combines with impurities to form slag.

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$O + SiO_2 \longrightarrow CaSiO_3$
Calcium silicate (slag)

- Coke, the source of chemical energy in the blast furnace, is burnt both to release heat energy and to provide the main reducing agent.

$$2C + O_2 \longrightarrow 2CO$$

— CO is used to reduce the oxide to molten iron

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2.$$

Chemical Reactions of Iron.

1) Rust formation; Rustling is the common term for corrosion of Iron and its alloys, such as steel. It is the most important reaction of iron from an economic point of view; essentially, rusting is the formation of hydrated iron (III) oxide in the presence of oxygen and water. The process is electrolytic;

- Water vapour on the metal surface dissolves CO_2 and O_2 from air forming carbonic acid which acts as the electrolytic solution of the cell.

$$H_2O + CO_2 \longrightarrow H_2CO_3.$$

- Iron, in contact with the dissolved CO_2 undergoes oxidation as follows:

Anodic reaction $Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^-$ (oxidation)

Cathodic reaction $O_2(g) + 2H_2O(l) + 4e \longrightarrow 4OH^-_{(aq)}$ (Reduction) overall reaction of cell,

$2Fe_{(s)} + O_{2(g)} + 2H_2O_{(l)} \rightarrow 2Fe^{2+}_{(aq)} + 4OH_{(aq)}$

The ferrous ion formed reacts with dissolved oxy:
to form ferric oxide:

$4Fe + O_2 + 4H_2O \rightarrow 2Fe_2O_3 + 8H^+_{(aq)}$

Ferric oxide then undergoes hydration to form Ru

$Fe_2O_3 + xH_2O \longrightarrow Fe_2O_3\,xH_2O$

The overall equation for rust formation is

$4Fe + 3O_2 + xH_2O \longrightarrow 2Fe_2O_3.xH_2O$

$Fe_2O_3.xH_2O$ (hydrated iron (III) oxide) is the flal
reddish - ~~brown~~ solid known as rust.

2) Reaction with water:
Fe reacts with steam to give iron oxide c
hydrogen gas

$3Fe_{(s)} + 4H_2O_{(g)} \longrightarrow Fe_3O_{4(s)} + 4H_{2(g)}$

3) Reaction with oxygen: strong heating is req
to make iron powder burn in oxygen to form
oxides

$4Fe_{(s)} + 3O_{2(g)} \longrightarrow 2Fe_2O_{3(s)}$

$3Fe_{(s)} + 2O_{2(g)} \longrightarrow Fe_3O_{4(s)}$

4) Reaction with acids.
Fe reacts with dilute acid to give the respe
salts and hydrogen gas.

$Fe_{(s)} + 2HCl_{(aq)} \longrightarrow FeCl_{2(aq)} + H_{2(g)}$

i) Reaction with the halogens

Iron reacts with excess of the halogens F_2, Cl_2 and Br_2, to form ferric, (Fe III), halides

$2Fe_{(s)} + 3Cl_{2(g)} \longrightarrow 2FeCl_{3(s)}$

Compounds of Iron

Iron forms two series of compounds which correspond to the ability of the atoms to form Iron (II) ion Fe^{2+} (ferrous compounds) and Iron (III) ion Fe^{3+} (ferric compounds). Iron II compounds are reducing agents, while Iron III compounds are oxidizing agents.

The following chemical tests can be used to distinguish between Iron II and Iron III compounds

i) reaction with NaOH or KOH solution: a solution of Iron II salt produces a pale green precipitate while that of Iron III salt produces a reddish brown precipitate.

ii) reaction with potassium hexacyanoferrate (III) solution K_3FeCN_6. Iron (II) solution gives a dark blue precipitate while Iron (III) salt gives a brown solution.

Examples of Iron (II) compounds are Iron II sulphate ($FeSO_4$) Iron II oxide (FeO), Iron III hydroxide

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Fe + NaOH → FeOH + Na

e$(OH)_2$), Iron II Chloride ($FeCl_2$)

Iron(III) Compounds are; Iron III oxide (Fe_2O_3), Iron III Chloride ($FeCl_3$), Iron III hydroxide $Fe(OH)_3$

Iron(II) Sulphate ($FeSO_4$) is used medically to treat Iron deficiency and in industrial applications such as Manufacture of ink and in dyeing. It is formed by the reaction of Iron with dilute H_2SO_4.

$$Fe + H_2SO_4 \longrightarrow FeSO_4 + H_2$$

Iron III Chloride ($FeCl_3$) is made by the direct Combination of the elements,

$$2Fe + 3Cl_2 \longrightarrow 2FeCl_3$$

Its major use is in Sewage treatment and as a Catalyst in organic synthesis.

Aluminium (Al)

Aluminium has many known isotopes, whose mass numbers range from 21 to 42, however, only ^{27}Al (stable isotope) and ^{25}Al radioactive isotope, occur naturally, ^{27}Al has a natural abundance above 99.9%

$FeSO_4$ FeO $FeCl_2$ $Fe(OH)_2$

Occurrence and Extraction

. The common ores of aluminium are: Bauxite ($Al_2O_3 \cdot nH_2O$) Corondum Al_2O_3, Kaolin ($Al_2O_3 \cdot 2SiO_2 \cdot 2H_2O$) Potash mica ($K_2O \, 3Al_2O_3 \, 6SiO_2 \, H_2O$) Cryolite ($Na_3AlF_6$)

Aluminium metal is extracted from bauxite in a two stage process as follows.

1) Purification of bauxite.

— bauxite is first roasted to eliminate water and then ground into powder

— the powdered bauxite is digested in conc NaOH to produce sodium dioxoaluminate III,

$$Al_2O_3 + 2NaOH \longrightarrow 2NaAlO_2 + H_2O$$

— The mixture is then filtered, the filtrate contains $NaAlO_2$ while the residue contains the impurities. The filtrate is diluted with water to produce aluminium hydroxide (a process referred to as 'seeding)

$$2NaAlO_2 + H_2O \longrightarrow Al(OH)_3 + 2NaOH$$

— the precipitated aluminium hydroxide is filtered off, dried and heated strongly to obtain pure aluminium oxide (alumina): $2Al(OH)_3 \longrightarrow Al_2O_3 + 3H_2O$

Continuations

2) Electrolysis of alumina (Al_2O_3):
Pure aluminium metal is produced by the electrolysis of alumina in molten cryolite. The aim of the cryolite is to lower the melting point in order for the electrolytic process to be carried out at a much lower temperature than if only Al_2O_3 is used.

Summary of electrolytic process;
Electrolyte ionizes thus: $Al_2O_3 \longrightarrow 2Al^{3+} + 3O^{2-}$

At the Anode: $6O^{2-} \longrightarrow 3O_2 + 12e^-$

At the Cathode $4Al^{3+} + 12e^- \longrightarrow 4Al$.

Chemical Reactivity of Al.

1) Corrosion resistance is excellent in aluminium due to a thin layer of aluminium oxide that forms when the metal is exposed to air, effectively preventing further oxidation

$$4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$$

The thin layer of oxide helps protect the metal from attack, However, if the oxide layer is damaged, the aluminium metal is exposed to attack.

2) Reaction with water: Al reacts with steam to form aluminium oxide and liberate hydrogen

$$2Al + 3H_2O \longrightarrow Al_2O_3 + 3H_2$$

3) Reaction with the halogens: Aluminium metal reacts vigorously with all the halogens to form aluminium halides.

$$4Al_{(s)} + 3Cl_{2(l)} \longrightarrow 2Al_2Cl_{3(s)}$$

4) Reaction with acids: Aluminium metal dissolves readily in dilute acids to form solutions containing the aquated Al(III) ion together with hydrogen gas.

$$2Al_{(s)} + 6HCl_{(aq)} \longrightarrow 2Al^{3+}_{(aq)} + 6Cl_{(aq)} + 3H_{2(g)}.$$

5) Reaction with mercury: In the presence of mercury, aluminium readily forms an amalgam (destroying the original surface) which is therefore, rapidly attacked by water, Contact with mercury Salts must be avoided if rapid Corrosion and weakening of aluminium structures is to be prevented.

Aluminium Compounds

Potassium Alum ($K_2SO_4Al_2(SO_4)_3 24H_2O$); Potassium alum is used as a coagulant in water purification processes and as a binding agent in the textile industry. It is prepared by mixing hot aqueous Solutions of K_2SO_4 and $Al_2(SO_4)_3$ in the ratio 1:1

$K_2SO_4 . Al_2(SO_4)_3 24H_2O$

On cooling, crystals of Potassium alum separate out.

Nitrogen (N)

Nitrogen is a gas which exists as a discrete diatomic molecules N_2. It is peculiar in showing a wide range of oxidation states eg +1, +5, −3. Nitrogen atoms in the Nitrogen molecule are held by a triple covalent bond ($N \equiv N$), the strength of which is shown by very high dissociation energy (946 KJ/mol), almost twice as large as that for an $O=O$ double bond

$$N_2(g) \longrightarrow 2N(g) \quad \Delta H = 946 \text{ KJ mol}^{-mol}$$

This makes it one of the strongest chemical bonds known and this is primarily responsible for the inertness of N_2.

Occurence, Isolation and Uses of Nitrogen

The most important source of nitrogen is the atmosphere, which consists of about 78% nitrogen by volume. Nitrogen is also found in combination with other elements in the earth crust.

Nitrogen is obtained industrially by the fractional distillation of liquified air. The major fractions are nitrogen, boiling point 77K and oxygen, boiling point 90K.

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Nitrogen Compounds

~~Nitrogen Compounds~~. Some important Compou of nitrogen are; Ammonia (NH_3), Trioxonitrate (V) acid (HNO_3), dinitrogen oxide or nitrous oxide, laughing gas), nitrogen (II) oxide (NO), nitrog oxide (NO_2).

1 Ammonia (NH_3) is manufactured by the Haber developed by a German Chemist fritz Haber i 1914. A mixture of nitrogen and hydrogen in ratio 1: 3 by volume is ~~s~~ubjected to a pressure 200 - 250 atm at about 450°c in the presence finely divided iron as Catalyst.

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

The reaction is exothermic and reversible.

2) Trioxonitrate (V) acid (HNO_3) is prepared in the om Potassium trioxonitrate V and tetraoxosu (VI) acid.

$$KNO_{3(s)} + H_2SO_{4(l)} \longrightarrow KHSO_{4(aq)} + HNO_{3(l)}$$

The major uses of HNO_3 are in the manufacture of fertilizers and drugs and as a nitrating agent in dye stuff.

Chemical Reactivity of Nitrogen

N_2 is so inert that lithium is one of the few elements with which it reacts with at room temperature.

$$6Li_{(s)} + N_2(g) \rightleftharpoons 2Li_3N(s)$$

Despite the fact that the N_2 molecule is unreactive, compounds containing nitrogen exist for virtually every element in the periodic table except those in Group VIIIA.

This can be explained in two ways. First, N_2 becomes significantly more reactive as the temperature increases.

Secondly, a number of catalysts found in nature help overcome the inertness of N_2 at low temperature

Reaction with Hydrogen; At high temperatures, nitrogen reacts with hydrogen to form ammonia.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

This reaction is the basis of the Haber process (for manufacture of ammonia)

Reaction with oxygen; Nitrogen reacts with oxygen at high temperature to form nitrogen oxide.

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$$

Nitrogen Compounds

~~Nitrogen Compounds~~: Some important Compounds of nitrogen are; Ammonia (NH_3), Trioxonitrate (V) acid (HNO_3), dinitrogen oxide or nitrous oxide, N_2O (laughing gas), nitrogen (II) oxide (NO), nitrogen IV oxide (NO_2).

i) Ammonia (NH_3) is manufactured by the Haber process (developed by a German Chemist fritz Haber in 1914. A mixture of nitrogen and hydrogen in the ratio 1:3 by volume is ~~S~~ subjected to a pressure of 200-250 atm at about 450°c in the presence of finely divided iron as Catilyst.

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

The reaction is exothermic and reversible.

ii) Trioxonitrate (V) acid (HNO_3) is prepared in the lab from potassium trioxonitrate V and tetraoxosulphate (VI) acid.

$$KNO_{3(s)} + H_2SO_{4(l)} \longrightarrow KHSO_{4(aq)} + HNO_{3(l)}$$

The major uses of HNO_3 are in the manufacture of fertilizers and drugs and as a nitrating agent in dye stuff.

Dinitrogen oxide, is known as laughing gas due to the euphoric effects of inhaling it. It is used in surgery and dentistry as an analgesic and anesthetic.
It is prepared in the lab by gently heating ammonium trioxonitrate (V) crystals;

$$NH_4NO_{3(s)} \longrightarrow 2H_2O + N_2O_{(g)}$$

Oxygen

Oxygen exists naturally as a gaseous diatomic molecule O_2. The two atoms are joined by a double covalent bond. The electron configuration of oxygen atom — $[He]\ 2s^2\ 2p^4$ — suggests that O atoms can achieve an octet of valence electrons by sharing two pairs of electrons. O_2 molecule is paramagnetic (it is attracted to a magnetic field). This can be explained using molecular orbital theory, which predicts that there are two unpaired electrons in the antibonding molecular orbitals of the O_2 molecule.

Occurrence, Isolation and preparation

Oxygen exists in two allotropic forms, Oxygen, O_2 and Ozone (trioxygen) O_3.

Oxygen O_2 is the most abundant element on this planet, it Constitutes about 21% by volume of the atmosphere. It is also found in Combination with other elements as mineral ores and also Constitute 90% by mass of water.

Ozone (O_3) is found in trace quantities in the upper atmosphere where it is believed to be formed by the photochemical dissociation of oxygen molecules by the intense ultra-violet light from the Sun.

$$O_{2(g)} + hv \longrightarrow 2O_{(g)}$$

$$[O] + O_2 \longrightarrow O_{3(g)}$$

Absorption of this light in the process prevents it from reaching the earth where it would destroy all living matter very rapidly. The Ozone layer is the earth's protective Shield.

Ozone is prepared in the laboratory by passing Silent electrical discharges through dry oxygen in an apparatus known as the Ozoniser. The commonly Used ozoniser is known as the Siemen's Ozoniser.

Preparation of oxygen

In the Laboratory, Oxygen may be prepared by different methods;

i) By the decomposition of hydrogen Peroxide, H_2O_2, a reaction Catalysed by Manganese (IV) oxide:

$$2H_2O_2 \longrightarrow 2H_2O + O_2\uparrow$$

ii) Thermal deComposition of oxides of heavy metals or Certain higher oxides;

$$2AgO \longrightarrow 2Ag + O_2\uparrow$$

$$2Pb_3O_4 \longrightarrow 6PbO + O_2\uparrow$$

iii) By the thermal deComposition of Salts Containing anions rich in oxygen

$$2KClO_3 \longrightarrow 2KCl + 3O_2$$

iv) Oxygen (O_2) is obtained in industrial scale by the fractional distillation of liquified air.

Chemical Reactivity of Oxygen

1) Oxygen as an oxidizing Agent: Oxygen is the Perfect example of an oxidizing agent because it increases the oxidation state of almost any Substance with which it reacts and gains electrons in virtually all of its Chemical reactions (it is reduced)