STUDY OF COMPOUNDS

HYDROGEN CHLORIDE

Hydrogen chloride was discovered in 1648. It is a very useful chemical compound. It occurs in free state in volcanic gases and in gastric juices of mammals. It was first prepared in 1648 by scientist called Johann Rudolf Glauber using rock salt and concentrated sulphuric acid. In 1772, Joseph Priestley was able to obtain it in pure form.

Molecular formula: HCl
Molecular mass: 36.5

Preparation methods:

Combination reaction:

Hydrogen chloride can be prepared by the reaction of moist hydrogen gas with chlorine in the presence of the diffused sunlight. The reaction should be maintained in diffused but not in direct sun light as it is explosive reaction. And not even in dark as it is very slow.
\[
\text{H}_2 + \text{Cl}_2 \quad \text{Diffused Sunlight} \rightarrow \quad 2\text{HCl (g)}
\]

\[
\text{H}_2 + \text{Cl}_2 \quad \text{Direct Sunlight} \rightarrow \quad 2\text{HCl (g) (Explosive)}
\]

\[
\text{H}_2 + \text{Cl}_2 \quad \text{Dark} \rightarrow \quad 2\text{HCl (g) (Negligible)}
\]

**Laboratory preparation of Hydrogen chloride:**

Reaction of sodium chloride with sulphuric acid:
Heating sodium chloride with concentrated sulphuric acid in a round-bottomed flask.

\[
\text{NaCl} + \text{H}_2\text{SO}_4 \quad 420\text{K} \rightarrow \quad \text{NaHSO}_4 + \text{HCl}
\]

Sodium chloride     Conc. Sulphuric acid     Sodium bisulphate     Hydrogen chloride

Hydrogen chloride can be prepared in laboratory by treating Sodium chloride with concentrated sulphuric acid.
During the preparation when the mixture of concentrated sulphuric
acid and sodium chloride which is in round bottom flask subjected to heating will produce hydrogen chloride gas. This hydrogen chloride is passed through a drying agent and is collected by upward displacement of air. Drying agent is used to remove moisture present in the product. Concentrated sulphuric acid (H\textsubscript{2}SO\textsubscript{4}), Phosphorus pentoxide (P\textsubscript{2}O\textsubscript{5}) and Calcium oxide (CaO) are commonly used drying agents.

**Confirmation of release of Hydrogen chloride:**

When inverted over a jar of ammonia gas kept near hydrogen chloride jar then dense white fumes are formed. These fumes consist of particles of solid ammonium chloride suspended in air. Combination of hydrogen chloride with ammonia forms solid ammonium chloride. Formation of white dense fumes confirms the formation of hydrogen chloride.

**Physical properties:**

- Hydrogen chloride is a colourless gas.
- It has pungent odour.
- It is slightly sour in taste.
- Fumes in moist air
- Not combustible
- Just extinguishes a glowing splint
- Very soluble in water; solution called hydrochloric acid
- Boiling point: –83ºC
- Freezing point: –113º C
Chemical properties:

Aqueous solution of hydrogen chloride is acidic in nature and changes blue litmus paper red. An aqueous solution reacts with metals, metal oxides, hydroxides, carbonates and bicarbonates to form the corresponding chlorides.

\[
2\text{HCl} + \text{Ca} \rightarrow \text{CaCl}_2 + \text{H}_2
\]

\[
2\text{HCl} + \text{MgO} \rightarrow \text{MgCl}_2 + \text{H}_2\text{O}
\]

\[
\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}
\]

\[
2\text{HCl} + \text{CaCO}_3 \rightarrow \text{CaCl}_2 + \text{H}_2\text{O} + \text{CO}_2
\]

\[
\text{HCl} + \text{NaHCO}_3 \rightarrow \text{NaCl} + \text{H}_2\text{O} + \text{CO}_2
\]

Formation of aqua regia:

Mixture of three parts of concentrated hydrogen chloride with one part of concentrated nitric acid is called aqua regia.

\[
3\text{HCl} + \text{HNO}_3 \rightarrow \text{NOCl} + 2\text{H}_2\text{O} + 2[\text{Cl}]
\]

The nascent chloride released during the mixing can react with noble metals like gold, platinum to form the metal chlorides.

Reaction with nitrates:

In general diluted HCl does not react with nitrates but can react with silver nitrate to form a white precipitate of silver chloride.

\[
\text{HCl} + \text{AgNO}_3 \rightarrow \text{AgCl} + \text{HNO}_3
\]

Uses:

Hydrogen chloride is useful to prepare aqua regia. It is useful to manufacture chlorine and some other chlorides. It is useful as a solvent for noble metals. It is useful as a reagent in laboratories. A saturated solution of zinc chloride in dilute hydrochloric acid is used to clean metallic surfaces before plating or soldering.
AMMONIA

Ammonia was earlier known by the name "alkaline air". It was identified in 1716 by J. Kunckel. It was prepared in a lab for the first time by Joseph Priestly in 1774. Although Claude Berthelot studied the chemical composition of the gas in 1785, it was Davy who proved in 1800 that it was composed of the elements nitrogen and hydrogen.

Molecular formula of ammonia: NH₃
Molecular mass: 17
In nature, ammonia occurs in both free as well as in combined state.
Preparation of ammonia:

By heating urea with soda lime.

\[
\text{NH}_2\text{CONH}_2 + 2\text{NaOH} \xrightarrow{\Delta} \text{Na}_2\text{CO}_3 + 2\text{NH}_3
\]

Urea Sodalime Sodium Ammonia carbonate

Hydrolysis of metal nitrides:

Magnesium nitride reacts with water to produce a suspension of magnesium hydroxide in water along with ammonia.

\[
\text{Mg}_3\text{N}_2 + 6\text{H}_2\text{O} \rightarrow 3\text{Mg(OH)}_2 + 2\text{NH}_3
\]

Calcium nitride, on hydrolysis with warm water, expels ammonia by leaving calcium hydroxide particles in water to form a suspension.

\[
\text{Ca}_3\text{N}_2 + 6\text{H}_2\text{O} \rightarrow 3\text{Ca(OH)}_2 + 2\text{NH}_3
\]

Aluminium nitride, on reaction with warm water, gives a precipitate of aluminium hydroxide along with ammonia.

\[
\text{AlN} + 3\text{H}_2\text{O} \rightarrow \text{Al(OH)}_3 + \text{NH}_3
\]

Laboratory preparation:

By heating ammonium chloride and slaked lime and then passing the ammonia formed through a drying tower filled with calcium oxide.
Industrial preparation:

By the direct union of nitrogen and hydrogen with Haber's process. When dry nitrogen and hydrogen mixed in 1:3 ratio at higher pressure and is then heated electrically in catalytic chamber forms ammonia gas. Ammonia gas and unreacted nitrogen and hydrogen passed through cooling to liquefy the ammonia but unreacted nitrogen and hydrogen will remain in gaseous state and circulated back into reaction chamber. The rate of formation ammonia increases by the use of iron catalyst.
Physical properties:

- Ammonia is a colourless gas with characteristic pungent odour.
- It is non-poisonous in nature but can have adverse effects on the respiratory system if inhaled in large quantities.
- It is lighter than air.
- Ammonia can easily liquefiable even at low temperatures.
- Boiling point of ammonia: –33.5° C
- Freezing point of ammonia: –77.7° C
- It is highly soluble in water.
- It causes irritating burning sensation in nasal passage.
- Ammonia makes eyes tearful.

\[ \text{N}_2 + 3\text{H}_2 \xrightarrow{\text{Fe}, \text{Mo}} 2\text{NH}_3 + \Delta \]
**Chemical properties:**

Ammonia gives reddish brown precipitate with Nessler's reagent. Neither combustible nor a supporter of combustion.

**Basic nature:**

In general, dry ammonia is neutral and aqueous solution of ammonia acts as a weak base.

**Reaction with acids:**

Ammonia can react with acids to form ammonium salts.
\[
\text{NH}_3 (g) + \text{HCl} (g) \rightarrow \text{NH}_4\text{Cl} (s)
\]
\[
\text{NH}_3 (g) + \text{H}_2\text{SO}_4 \rightarrow (\text{NH}_4)_2\text{SO}_4
\]

**Reaction with metallic salts:**

Ammonia dissolved in water to form ammonium hydroxide. Ammonium hydroxide reacts with metallic salts to form insoluble precipitates.
\[
\text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_4\text{OH}
\]
\[
\text{NH}_4\text{OH} + \text{FeCl}_3 \rightarrow 3\text{NH}_4\text{Cl} + \text{Fe(OH)}_3
\]
\[
\text{NH}_4\text{OH} + \text{CuSO}_4 \rightarrow (\text{NH}_4)_2\text{SO}_4 + \text{Cu(OH)}_2
\]

**Formation of urea:**

Ammonia can react with carbon dioxide to form urea at 423K temperature and 150 atmospheric pressure.
\[
2\text{NH}_3 + \text{CO}_2 \rightarrow \text{NH}_2\text{CONH}_2 + \text{H}_2\text{O}
\]
**Dissociation of ammonia:**

Can dissociate to form nitrogen and hydrogen when passed through a red-hot tube or electric spark.

\[ 2\text{NH}_3 \rightleftharpoons \text{N}_2 + 3\text{H}_2 \]

**Oxidation of ammonia:**

Ammonia reacts with oxygen when heated in the presence of platinum at 800° C to yield nitric oxide.

\[ 4\text{NH}_3 + 5\text{O}_2 \rightarrow 4\text{NO} + 6\text{H}_2\text{O} + \text{heat energy} \]
\[ 2\text{NO} + \text{O}_2 \rightarrow 2\text{NO}_2 \]

**Reaction with metal oxides:**

Ammonia reduces metal oxides to free metals

\[ 2\text{NH}_3 + 3\text{CuO} \rightarrow 3\text{Cu} + 3\text{H}_2\text{O} + \text{N}_2 \]

**Reduction of halogens:**

Ammonia reduces halogens to hydrogen halides.

\[ 2\text{NH}_3 + 3\text{Cl}_2 \rightarrow \text{N}_2 + 6\text{HCl} \]

**Uses:**

Ammonia is a popular laboratory reagent

It is useful to clean grease.

Ammonia can be useful as a refrigerant in an ice plant.

Ammonia is useful in the manufacture of washing soda and baking soda, Urea, ammonium phosphate, ammonium sulphate.
It is also useful manufacture of plastics, explosives, artificial silk, nitric acid, dyes and drugs.

**NITRIC ACID**

Nitrogen forms a number of oxoacids, such as hypo nitrous acid, nitrous acid, per nitric acid and nitric acid. Of these, nitric acid is an industrially important oxoacid.

**Methods to prepare nitric acid:**

In the laboratory, nitric acid is prepared by heating potassium nitrate or sodium nitrate with concentrated sulphuric acid.

\[
\text{KNO}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{KHSO}_4 + \text{HNO}_3
\]

potassium nitrate sulphuric acid Nitric acid

\[
\text{NaN}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{NaHSO}_4 + \text{HNO}_3
\]

Sodium nitrate sulphuric acid Nitric acid

On a commercial scale, it is prepared by Ostwald’s process. The first step in this process is the catalytic oxidation of ammonia into nitric oxide.

\[
4\text{NH}_3(g) + 5\text{O}_2(g) + 6\text{H}_2\text{O}(g) \rightarrow 4\text{NO}(g)
\]

potassium nitrate sulphuric acid 1155 K, 9 bar Nitric oxide

The nitric oxide is then oxidised to nitrogen dioxide, from which nitric acid is formed.

\[
2\text{NO}(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g)
\]
Nitric oxide + Oxygen → Nitrogen dioxide
3NO₂(g) + H₂O(g) → 2HNO₃(aq) + NO(g)
Nitrogen dioxide + Water → Nitric acid

98% concentrated nitric acid, called fuming nitric acid.

**Properties of nitric acid:**

It is a colourless liquid that boils at 84.1°C and freezes to a white solid at -41.55°C.

Nitric acid undergoes photochemical dissociation to produce brown nitrogen dioxide gas. The brown nitrogen dioxide dissolves in the colourless nitric acid, turning it yellow.

\[
\text{Sunlight} \quad 4\text{HNO}_3 \quad \rightarrow \quad 4\text{NO}_2 \quad + \quad 2\text{H}_2\text{O} \\
\text{Nitric acid} \quad \text{Nitrogen dioxide}
\]

An aqueous solution of nitric acid undergoes ionisation to produce hydronium ions and nitrate ions.

\[
\text{HNO}_3(\text{aq}) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(\text{aq}) \quad + \quad \text{NO}_3^-(\text{aq}) \\
\text{Nitric acid} \quad \text{Hydronium ions} \quad \text{Nitrate ions}
\]

It is a strong acid. Hot and concentrated nitric acid acts as a powerful oxidising agent. Metals such as copper and zinc are insoluble in hydrochloric acid. However, these metals dissolve in nitric acid. Concentrated nitric acid reacts with copper and zinc to give nitrogen dioxide.

\[
\text{Cu} \quad + \quad 4\text{HNO}_3 \quad \rightarrow \quad \text{Cu(NO}_3)_2 \quad + \quad 2\text{NO}_2 \quad + \quad 2\text{H}_2\text{O}
\]
Copper Nitric acid (Conc.) Copper nitrate Nitrogen dioxide

\[
\text{Zn} + 4\text{HNO}_3 \rightarrow \text{Zn(NO}_3)_2 + 2\text{NO}_2 + 2\text{H}_2\text{O}
\]
Zinc Nitric acid (Conc.) Zinc nitrate Nitrogen dioxide

Dilute nitric acid reacts with copper to give nitric oxide, and with zinc to give nitrous oxide.

\[
3\text{Cu} + 8\text{HNO}_3 \rightarrow 3\text{Cu(NO}_3)_2 + 2\text{NO} + 4\text{H}_2\text{O}
\]
Copper Nitric acid (Dilute) Copper nitrate Nitric oxide

\[
4\text{Cu} + 10\text{HNO}_3 \rightarrow 4\text{Zn(NO}_3)_2 + 2\text{N}_2\text{O} + 5\text{H}_2\text{O}
\]
Zinc Nitric acid (Dilute) Zinc nitrate Nitric oxide

Certain metals such as chromium, aluminium and iron are rendered passive towards an attack by nitric acid. This is due to the formation of a thin layer of oxide on the surface of these metals.

**Aqua regia:**

It is the mixture of 1 part of conc. HNO₃ & 3 parts of conc. HCl. Noble metals like gold and platinum which cannot dissolve in conc. HNO₃ can be dissolved in aqua regia. Concentrated nitric acid oxidises non-metals such as carbon, phosphorus, sulphur and iodine to their corresponding higher oxoacids. The nascent oxygen produced from the dissociation of nitric acid oxidises these non-metals.

\[
\text{C} + 4\text{HNO}_3 \rightarrow \text{H}_2\text{CO}_3 + 2\text{H}_2\text{O} + 4\text{NO}_2
\]
Carbon Nitric acid Carbonic acid Nitrogen
dioxide

\[ P_4 + 20\text{HNO}_3 \rightarrow 4\text{H}_3\text{PO}_4 + 4\text{H}_2\text{O} + 20\text{NO}_2 \]

Carbon dioxid \hspace{1cm} Nitric acid \hspace{1cm} Phosphoric acid \hspace{1cm} Nitrogen dioxide

\[ S_8 + 48\text{HNO}_3 \rightarrow 8\text{H}_2\text{SO}_4 + 16\text{H}_2\text{O} + 48\text{NO}_2 \]

Sulphur dioxid \hspace{1cm} Nitric acid \hspace{1cm} Sulphuric acid \hspace{1cm} Nitrogen dioxide

To confirms the presence of nitrate ions in a given salt brown ring test conducted.

**Uses of nitric acid:**

- Nitric acid is widely used in the manufacture of dyes and drugs, as well as explosives such as trinitrotoluene, or TNT, nitroglycerine and picric acid.
- The salts of nitric acid, such as ammonium nitrate and calcium nitrate, are used as important fertilisers.
- It is used in the purification of silver and gold.
- It is used in the laboratory as an oxidising and a nitrating reagent.
- It finds a major use in the pickling of stainless steel, etching of metals and as an oxidiser in rocket fuels.