

GCE O Level Pure Chemistry

Summary of Chemical Reactions & Household Chemicals

<http://www.jce.divched.org/JCESoft/CCA/CCA3/CONTENTS.HTM>

View the chemical reactions in Live Video from the above link

(A) Part-1 Important exception reactions by each types of chemical Reaction

I) Acid/Base Reactions

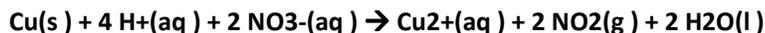


- Acid + Metal \rightarrow Salt + H₂
- Acid + Carbonate \rightarrow Salt + H₂O + CO₂
- Acid + Alkaline \rightarrow Salt + H₂O
- Acid + Base \rightarrow Salt + H₂O
- Conc Acid + Metal \rightarrow Salt + H₂O + gas

Of the four industrially important strong acids (HCl, H₃PO₄, H₂SO₄, HNO₃), only two, nitric and sulfuric, are strong enough oxidizing agents to dissolve copper. The concentration of copper(II) ions is larger with the nitric acid, because it is a stronger oxidizing agent than sulfuric acid.



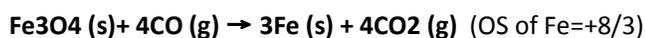
Concentrated sulfuric acid, formerly called oil of vitriol, is a weak acid and a poor electrolyte because relatively little of it is dissociated into ions. When concentrated sulfuric acid is heated, it behaves also as an oxidizing agent dissolving relatively unreactive metals as copper, mercury, and lead to produce metal sulfate, sulfur dioxide, and water



Red-brown nitrogen dioxide is generated and some of the copper dissolves to form a blue solution of copper(II) nitrate.

II) Redox Reactions

- The oxidation state can be fractional or whole number

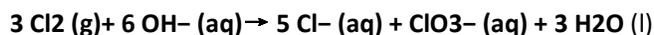


Iron(II,III) oxide is the chemical compound with formula Fe_3O_4 . It is one of a number of iron oxides, the others being iron(II) oxide (FeO), which is rare, and iron(III) oxide (Fe_2O_3) also known as **hematite**. It occurs in nature as the mineral [magnetite](#). It contains both Fe^{2+} and Fe^{3+} ions and is sometimes formulated as $\text{FeO} \cdot \text{Fe}_2\text{O}_3$.

This iron oxide is encountered in the laboratory as a black powder. It exhibits permanent magnetism and is ferrimagnetic, but is sometimes incorrectly described as ferromagnetic. Its most extensive use is as a black pigment which is synthesised rather than being extracted from the naturally occurring mineral as the particle size and shape can be varied by the method of production.

- Disproportionation, also known as dismutation is a specific type of redox reaction in which a species is simultaneously reduced and oxidized to form two different products.

Chlorine gas reacts with dilute sodium hydroxide to form sodium chloride, sodium chlorate and water. The ionic equation for this reaction is as follows :

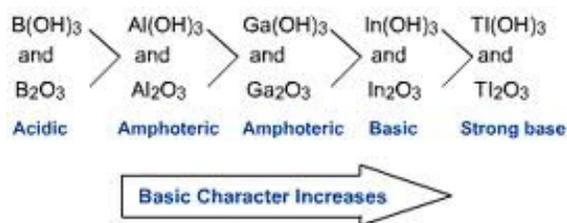


The chlorine gas reactant is in oxidation state 0. In the products, the chlorine in the Cl^- ion has an oxidation number of -1 , having been reduced, whereas the oxidation number of the chlorine in the ClO_3^- ion is $+5$, indicating that it has been oxidized.

The disproportionation of hydrogen peroxide into water and oxygen catalysed by either potassium iodide or the enzyme catalase:



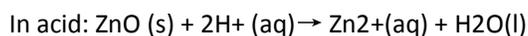
III) Amphoteric reactions



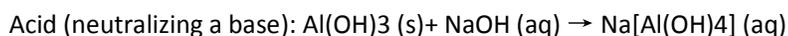
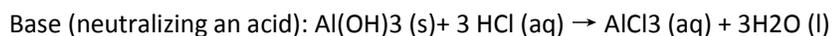
Amphoteric oxides and hydroxides

an amphoteric species is a molecule or ion that can react as an acid as well as a base. Many metals (such as zinc, tin, lead, aluminum, and beryllium) and most metalloids form amphoteric oxides or hydroxides.

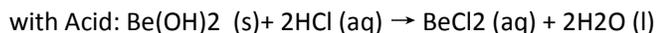
- Zinc oxide (ZnO)



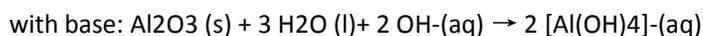
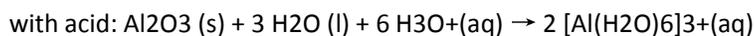
- Aluminum hydroxide



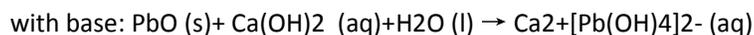
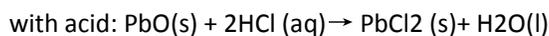
- Beryllium hydroxide



- Aluminum oxide



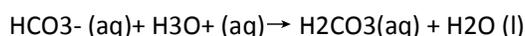
- Lead oxide



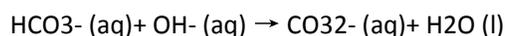
Some other elements which form amphoteric oxides are chromium, gallium, copper, antimony, bismuth, indium, silicon, titanium, vanadium, iron, cobalt, germanium, zirconium, silver, tin, and gold.

Amphiprotic molecules

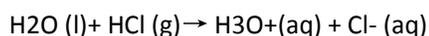
- hydrogen carbonate ion, which can act as a base:



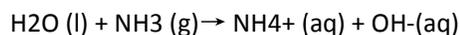
Or as an acid:



- Water is the most common example, acting as a base when reacting with an acid such as hydrogen chloride

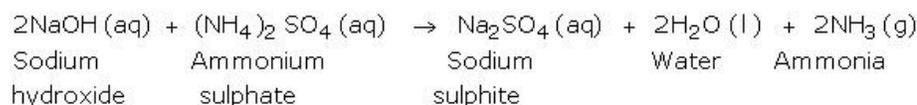
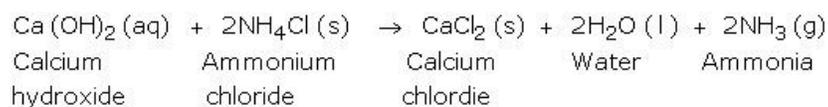


and acting as an acid when reacting with a base such as ammonia:



IV) Alkaline reaction with Ammonia salt

Alkali + Ammonium salt \rightarrow Salt + Water + Ammonia



V) Reversible reactions

A reversible reaction is a chemical reaction that results in an equilibrium mixture of reactants and products. For a reaction involving two reactants and two products this can be expressed symbolically as



A truly irreversible chemical reaction is usually achieved when one of the products exits the reacting system, for example, as carbon dioxide (volatile) in the reaction



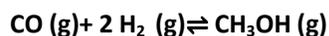
Le Chatelier's principle can be used to predict the effect of a change in conditions on a chemical equilibrium. It can be summarized as:

If a chemical system at equilibrium experiences a change in concentration, temperature, volume, or partial pressure, then the equilibrium shifts to counteract the imposed change and a new equilibrium is established.

Effect of change in concentration

Changing the concentration of an ingredient will shift the equilibrium to the side that would reduce that change in concentration. The chemical system will attempt to partially oppose the change affected to the original state of equilibrium. In turn, the rate of reaction, extent and yield of products will be altered corresponding to the impact on the system.

This can be illustrated by the equilibrium of carbon monoxide and hydrogen gas, reacting to form methanol.

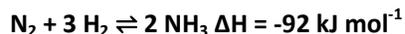


Suppose we were to increase the concentration of CO in the system. Using Le Chatelier's principle, we can predict that the amount of methanol will increase, decreasing the total change in CO. As the concentration of CO is increased, the frequency of successful collisions of that reactant would increase also, allowing for an increase in forward reaction, and generation of the product. Even if a desired product is not thermodynamically favored, the end-product can be obtained if it is continuously removed from the solution.

Effect of change in temperature

The effect of changing the temperature in the equilibrium can be made clear by incorporating heat as either a reactant or a product. When the reaction is exothermic (ΔH is negative, puts energy out), we include heat as a product, and, when the reaction is endothermic (ΔH is positive, takes energy in), we include it as a reactant. Hence, we can determine whether increasing or decreasing the temperature would favour the forward or reverse reaction by applying the same principle as with concentration changes.

For example, the reaction of nitrogen gas with hydrogen gas. This is a reversible reaction, in which the two gases react to form ammonia:

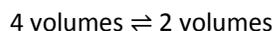
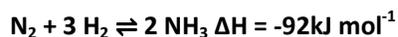


This is an exothermic reaction (hence the minus sign) when producing ammonia. If we were to lower the temperature, the equilibrium would shift to produce more heat. Since making ammonia is exothermic, this would favour the production of more ammonia. In practice, in the Haber process, the temperature is set at a compromise value, so ammonia is made quickly, even though less would be present at equilibrium.

Effect of change in pressure

Changes in pressure are attributable to changes in volume. The equilibrium concentrations of the products and reactants do not directly depend on the pressure subjected to the system. However, a change in pressure due to a change in volume of the system will shift the equilibrium.

Considering the reaction of nitrogen gas with hydrogen gas to form ammonia:



Note the number of moles of gas on the left-hand side and the number of moles of gas on the right-hand side. When the volume of the system is changed, the partial pressures of the gases change. If we were to decrease pressure by increasing volume, the equilibrium of the above reaction will shift to the left, because the reactant side has greater number of moles than does the product side. The system tries to counteract the decrease in partial pressure of gas molecules by shifting to the side that exerts greater pressure. Similarly, if we were to increase pressure by decreasing volume, the equilibrium shifts to the right, counteracting the pressure increase by shifting to the side with fewer moles of gas that exert less pressure. If the volume is increased because there are more moles of gas on the reactant side, this change is more significant in the denominator of the equilibrium constant expression, causing a shift in equilibrium.

Thus, an increase in system pressure due to decreasing volume causes the reaction to shift to the side with the fewer moles of gas. A decrease in pressure due to increasing volume causes the reaction to shift to the side with more moles of gas. There is no effect on a reaction where the number of moles of gas is the same on each side of the chemical equation.

Effect of adding an inert gas

An inert gas (or noble gas) such as helium is one that does not react with other elements or compounds. Adding an inert gas into a gas-phase equilibrium at constant volume does not result in a shift. This is because the addition of a non-reactive gas does not change the partial pressures of the other gases in the container. While it is true that the total pressure of the system increases, the total pressure does not have any effect on the equilibrium constant; rather, it is a change in partial pressures that will cause a shift in the equilibrium. If, however, the volume is allowed to increase in the process, the partial pressures of all gases would be decreased resulting in a shift towards the side with the greater number of moles of gas.

Effect of a catalyst

A catalyst speeds up the rate of a reaction by providing additional mechanism(s). Adding a catalyst allows for alternative pathways to be made, where the particles can be adsorbed onto the catalyst temporarily before being re-bonded into a new arrangement. The intended effect in adding a catalyst is to lower the activation energy, which frequently increases the rate of reaction. However, the activation energy is lowered by the same amount for the forward and reverse reactions. There is the same increase in reaction rates for both reactions. *As a result, a catalyst does not affect the position of the equilibrium. It only affects the time or energy that is required to achieve equilibrium.*

VI) Thermal Chemistry Bonds Strength

The Bond Enthalpy is the energy required to break a chemical bond. It is usually expressed in units of kJ mol^{-1} , measured at 298 K. *The exact bond enthalpy of a particular chemical bond depends upon the molecular environment in which the bond exists. Therefore, bond enthalpy values given in chemical data books are averaged values.*

Average Bond Enthalpies (kJ/mol)

Single Bonds

C—H	413	N—H	391	O—H	463	F—F	155
C—C	348	N—N	163	O—O	146		
C—N	293	N—O	201	O—F	190	Cl—F	253
C—O	358	N—F	272	O—Cl	203	Cl—Cl	242
C—F	485	N—Cl	200	O—I	234		
C—Cl	328	N—Br	243			Br—F	237
C—Br	276			S—H	339	Br—Cl	218
C—I	240	H—H	436	S—F	327	Br—Br	193
C—S	259	H—F	567	S—Cl	253		
		H—Cl	431	S—Br	218	I—Cl	208
Si—H	323	H—Br	366	S—S	266	I—Br	175
Si—Si	226	H—I	299			I—I	151
Si—C	301						
Si—O	368						

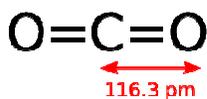
Multiple Bonds

C=C	614	N=N	418	O ₂	495
C≡C	839	N≡N	941		
C=N	615			S=O	523
C≡N	891			S=S	418
C=O	799				
C≡O	1072				

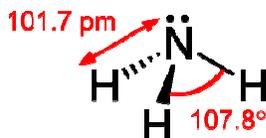
Bond diagrams of common gases:



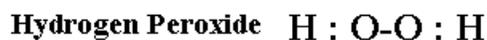
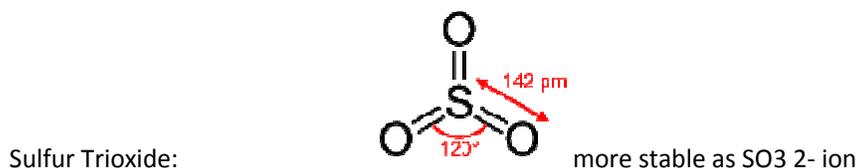
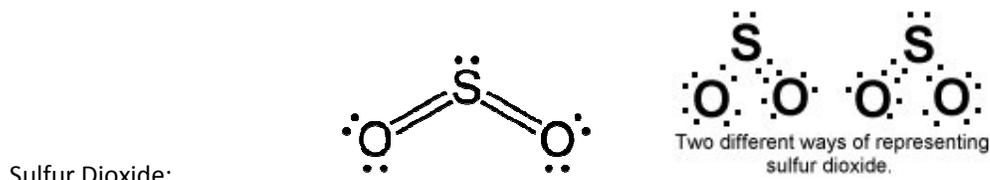
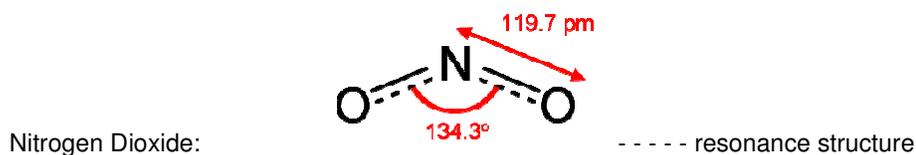
Carbon Monoxide:



Carbon Dioxide:



Ammonia:

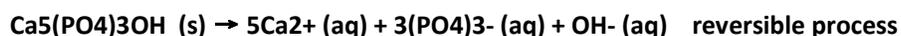


(B) Part-2 Chemistry Reactions in real life situation

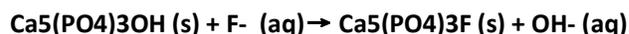
1) Tooth paste chemistry



Acid Base Reaction



Fluoride ion acts on the basic component of enamel, hydroxyapatite, $\text{Ca}_5(\text{PO}_4)_3\text{OH}$, by replacing the hydroxide ion OH^- . In this way, hydroxyapatite is replaced partially by the harder fluoro analogue fluoroapatite :



The most common active ingredients in toothpastes are :

- **Fluoride** is the most popular active ingredient in toothpaste due to its proved ability to prevent cavities. Most toothpaste brands use Sodium fluoride (NaF); some brands use Sodium Monofluorophosphate - SMFP ($\text{Na}_2\text{PO}_3\text{F}$).
- **Antimicrobial agents** that fight the bacteria of dental plaque.
- **Surfactants** (detergents) and Foaming agents that help to carry away debris from the mouth and between the teeth. Common foaming ingredients in toothpastes are Sodium Lauryl Sulfate (SLS) and ammonium lauryl sulfate.
- **Anti-tartar agents** as Tetrasodium Pyrophosphate (TSPP). Pyrophosphates are water-softening agents that remove calcium and magnesium from the saliva, so they can't deposit on teeth creating dental tartar (calcified plaque). Pyrophosphate does not remove tartar, it merely helps prevent its formation.
- **De-sensitizing agents** to relieve tooth sensitivity. Strontium chloride works by blocking the tiny crevices (microtubules) that enable cold and heat sensations to reach the tooth's nerve. Potassium citrate and Potassium nitrate work in a different way by blocking the mechanism of pain transmission between nerve cells.
- **Abrasives** : Calcium phosphate (chalk) and alumina (Al_2O_3) were used as the abrasive base of tooth pastes but they had the disadvantage of reacting with other chemical ingredients. Today the common abrasives are Silicon Dioxide (silica SiO_2) and Titanium Dioxide. Hydrated silica is a transparent abrasive used not only in white opaque tooth pastes, but in gel toothpastes as well. Baking Soda (Sodium bicarbonate) is a mild abrasive. It has a mild

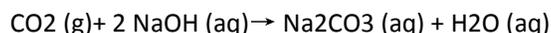
whitening action and helps to keep an alkaline environment (not friendly for dental plaque bacteria) in the mouth.

- **Teeth Whitening agents** : the whitening toothpastes, except of the mechanical whitening action of toothpaste abrasives, use extra whitening ingredients as the Hydrogen Peroxide or Sodium carbonate peroxide that breaks down into sodium carbonate (washing soda) and hydrogen peroxide.
- **Flavor.** Toothpastes come in a variety of flavors, most often being some variation on mint.
- Additional ingredients such as enzymes, vitamins, herbs, calcium, mouthwash are often included in the formulas. Other non-active ingredients in toothpaste are humectants, coloring thickeners, water softeners and sweeteners as sodium saccharin

2) Sodium bicarbonates (Baking Soda)



NaHCO₃ may be obtained by the reaction of carbon dioxide with an aqueous solution of sodium hydroxide. The initial reaction produces sodium carbonate:



Further addition of carbon dioxide produces sodium bicarbonate, which at sufficiently high concentration will precipitate out of solution:

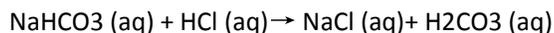


Acid/Base reaction:

Sodium bicarbonate is an amphoteric compound. Aqueous solutions are mildly alkaline due to the formation of carbonic acid and hydroxide ion:



Sodium bicarbonate can be used as a wash to remove any acidic impurities from a "crude" liquid, producing a purer sample. Reaction of sodium bicarbonate and an acid produce a salt and carbonic acid, which readily decomposes to carbon dioxide and water:

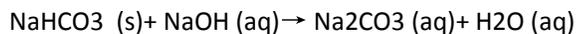




Sodium bicarbonate reacts with acetic acid (found in vinegar), producing sodium acetate, water, and carbon dioxide: (Many medicine and fruit salt uses this reaction to improvement dissolution and mixing of drugs in water)



Sodium bicarbonate reacts with bases such as sodium hydroxide to form carbonates:



Thermal decomposition

Above 70 °C, sodium bicarbonate gradually decomposes into sodium carbonate, water and carbon dioxide. The conversion is fast at 200 °C:[9]



Most bicarbonates undergo this dehydration reaction. Further heating converts the carbonate into the oxide (at ca. 1000 °C):



These conversions are relevant to the use of NaHCO₃ as a fire-suppression agent ("BC powder") in some dry powder fire extinguishers.

Applications

For the Home:

- Cleaning agent - fridges, garbage disposals, etc.
- Water softener – use in washing machines and for washing dishes
- Deodorizing - shoes, footwear cupboards, carpet refresher
- Fire extinguisher – forms a smothering soapy foam
- Polishing - removes dirt without damaging high hardness materials such as stainless steel or iron
- Mixed with sugar - acts as an effective pesticide for roaches and silverfish

In the Kitchen:

- Baking – used as a leavening agent, the carbon dioxide generated makes the dough rise
- Cooking – quickly softens vegetables such as french beans and broccoli

For Care of the Body:

- Use as a teeth whitener - brush teeth with a paste of bicarb soda and water

- Sodium bicarbonate ear drops
- Stomach reliever - spoonful in cool glass of water
- Eczema reliever - half-cup in hot bath from 15 to 20 minutes
- Also good for the throat, painful gums, insect bites, and warts
- Use in cosmetics and personal care products

For Emergencies:

Sodium bicarbonate can neutralize or reduce acids in the blood, or urine. It may be used in emergency medical situations (heart attacks, serious kidney or lung problems) to correct the normal acid-base balance in the blood or as an aid in treating overdoses with certain types of medications.

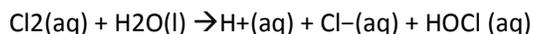
(3) Bleaching



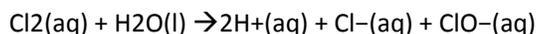
Bleach refers to a number of chemicals that remove color, whiten, or disinfect, often via oxidation. Common chemical bleaches include household chlorine bleach (a solution of approximately 3–6% sodium hypochlorite), lye, oxygen bleach (which contains either hydrogen peroxide or a peroxide-releasing compound), and bleaching powder (calcium hypochlorite).

Redox Reaction:

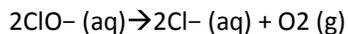
The active ingredient in bleach, hypochlorite ion, is produced by the following set of chemical reactions:



The H^+ ion of the hypochlorous acid then dissolves into solution, and so the final result is effectively:



Hypochlorite tends to decompose into chloride and a highly reactive form of oxygen:



Mechanism of bleach action

- **Whitening**

Colour in most dyes and pigments is produced by molecules which contain chromophores, such as beta carotene. Chemical bleaches work in one of two ways:

An oxidizing bleach works by breaking the chemical bonds that make up the chromophore. This changes the molecule into a different substance that either does not contain a chromophore, or contains a chromophore that does not absorb visible light.

A reducing bleach works by converting double bonds in the chromophore into single bonds. This eliminates the ability of the chromophore to absorb visible light.

Sunlight acts as a bleach through a process leading to similar results: high energy photons of light, often in the violet or ultraviolet range, can disrupt the bonds in the chromophore, rendering the resulting substance colorless. Extended exposure often leads to massive discoloration usually reducing the colors to white and typically very faded blue spectrums.

- **Color Safe Bleach**

Color safe bleach is a chemical that uses hydrogen peroxide as the active ingredient (to help remove stains) rather than sodium hypochlorite or chlorine. It also has chemicals in it that helps brighten colors. Hydrogen peroxide is also used for sterilization purposes and water treatment, but its disinfectant capabilities may be limited due to the concentration in the colorsafe bleach solution as compared to other applications. Because color safe bleach is mainly used to remove stains from color clothing while brightening colors, it is in a way the opposite of regular bleach. In fact, mixing the two together can cause them to cancel each other out some what (while the presence of the regular bleach could still ruin the regular color clothes though not as badly as with regular bleach alone)

(4) Household Ammonia



One of the most characteristic properties of ammonia is its basicity. It combines with acids to form salts; thus with hydrochloric acid it forms ammonium chloride. Dilute aqueous ammonia can be applied on the skin to lessen the effects of acidic animal venoms, such as from insects and jellyfish. The basicity of ammonia also is the basis of its toxicity and its use as a cleaner.

By creating a solution with a pH much higher than a neutral water solution, proteins (enzymes) will denature, leading to cell damage, death of the cell, and eventually death of the organism. Dirt often consists of fats and oils, which are sparingly soluble in water. Ammonia brings them into aqueous solution. The remaining water, also containing excess ammonia, will evaporate completely, leaving a clean surface.

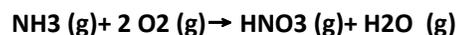
Uses

- Fertilizer

Approximately 83% (as of 2004) of ammonia is used as fertilizers either as its salts or as solutions. When applied to soil, it helps provide increased yields of crops such as corn and wheat. Consuming more than 1% of all man-made power, the production of ammonia is a significant component of the world energy budget.

- Precursor to nitrogenous compounds

Ammonia is directly or indirectly the precursor to most nitrogen-containing compounds. Virtually all synthetic nitrogen compounds are derived from ammonia. An important derivative is nitric acid. This key material is generated via the Ostwald process by oxidation of ammonia with air over a platinum catalyst at 700–850 °C, ~9 atm. Nitric oxide is an intermediate in this conversion:



Nitric acid is used for the production of fertilizers, explosives, and many organonitrogen compounds.

- Cleaner

Household ammonia is a solution of NH_3 in water (i.e., ammonium hydroxide) used as a general purpose cleaner for many surfaces. Because ammonia results in a relatively streak-free shine, one of its most common uses is to clean glass, porcelain and stainless steel. It is also frequently used for cleaning ovens and soaking items to loosen baked-on grime. Household ammonia ranges in concentration from 5 to 10 weight percent ammonia.

- Fermentation

Solutions of ammonia ranging from 16% to 25% are used in the fermentation industry as a source of nitrogen for microorganisms and to adjust pH during fermentation

(5) Descaling/Pickling of metal



Pickling is a metal surface treatment used to remove impurities, such as stains, inorganic contaminants, rust or scale from ferrous metals, copper, and aluminum alloys. A solution called pickle liquor, which contains strong acids, is used to remove the surface impurities. It is commonly used to descale or clean steel in various steelmaking processes.

Many hot working processes and other processes that occur at high temperatures leave a discoloring oxide layer or scale on the surface. In order to remove the scale the workpiece is dipped into a vat of pickle liquor.

The primary acid used is hydrochloric acid, although sulfuric acid was previously more common. Hydrochloric acid is more expensive than sulfuric acid, but it pickles much faster while minimizing base metal loss by using inhibitor.

Carbon steels, with an alloy content less than or equal to 6%, are often pickled in hydrochloric or sulfuric acid. Steels with an alloy content greater than 6% must be pickled in two steps and other acids are used, such as phosphoric, nitric and hydrofluoric acid. Rust- and acid-resistant chromium-nickel steels are pickled in a bath of hydrochloric and nitric acid. Most copper alloys are pickled in dilute sulfuric acid, but brass is pickled in concentrated sulfuric and nitric acid mixed with sodium chloride and soot.

By submerging steel in acid, two main reactions take place. Both use hydrogen chloride and both produce ferrous chloride.



The first reaction is where the acid reacts with the base metal (the steel under the scale). This reaction is quite slow and produces hydrogen gas as a byproduct, which accounts for the bubbling and foaming in the tank.

The second reaction is where the acid reacts with the scale itself. This reaction is much faster than the first and it produces water as a byproduct.

Iron Loss and Inhibitors

One of the most important criteria in a pickling and acid reclamation operation is the iron loss. Fortunately, some very effective materials called inhibitors are available to reduce the attack on the metal without affecting the scale removal. Inhibitors are very effective organic compounds which are added in low dosages, typically 0.1 to 0.3% by volume the pickling acid. The inhibitor attaches itself as a film to the clean steel and prevents further acid attack; the inhibitor does not adhere to the scale.

(6) Plastics



Common plastics and uses

Thermoplastics include:

Name	Properties	Applications and Uses
ABS (acrylonitrile butadiene styrene)	Strong, tough, hard, chemical resistant, durable. Suitable for all moulding processes.	Dashboards and car trim, toys, phones, handles, electrical products.
Nylon (or polyamide)	Tough, hard, light, self lubricating, chemical resistant, machines well, extrudable, injects well.	Bearings, gears, rope, hinges and catches, engineering applications.
Acetate (or cellulose)	Tough, stiff, hard, transparent, light, heat resistant.	Tool handles, pen bodies, frames for glasses. Can be injection moulded.
Acrylic (polymethylacrylate) Perspex (tradename) (polymethylmethacrylate, PMMA)	Stiff, durable, insulator, machines well, polishes well, scratches easily.	Car light covers, baths, shower trays, basins. Can be line bent/vacuum formed/injection moulded with ease.
Polypropylene (or polypropene)	Tough, light, chemical resistant. Will scratch, quite soft.	Containers, pots, plastic seats, ropes, nets. Very versatile.
Polystyrene expanded- high impact-	Light, stiff, transparent, brittle, waterproof/resistant. very light very tough	toys, electrical product cases, boxes packaging vacuum formings
Polythene (or polyethene) Low density (LDPE) High density (HDPE)	Insulator, chemical resistant. Tough, flexible, soft. Hard, stiff.	Packaging, bags, Bottles, domestic appliances.
PVC (polyvinyl chloride) Polychloroethene uPVC (U = unplasticised)	Stiff, hard, tough, light White	Cables, hoses, sheet, fabric Gutters, windows, extrusions

Thermoset plastics differ in that they are not re-mouldable. Strong cross links are formed during the initial moulding process that give the material a stable structure. They are more likely to be used in situations where thermal stability is required. They tend to lack tensile strength and can be brittle.

Name	Properties	Applications and Uses
Urea formaldehyde	Strong, insulator, brittle, hard, stiff.	Electrical fittings, handles and knobs
Polyester resin	Liquid raw state, stiff hard, insulator, chemical resistance, brittle without fibre reinforcement.	Casting, bonding fibres (glass, Kevlar, carbon fibre)
Epoxy resins trade names include Araldite.	Good insulator, brittle chemical resistant.	Adhesives, bonding fibres, encapsulation.
Melamine formaldehyde	Hard, strong, heat resistant.	Adhesives, bonding fibres, encapsulation.
Polycarbonate	Hard, strong, transparent, high refractive index	Spectacle lenses

(7) Carbonated Water



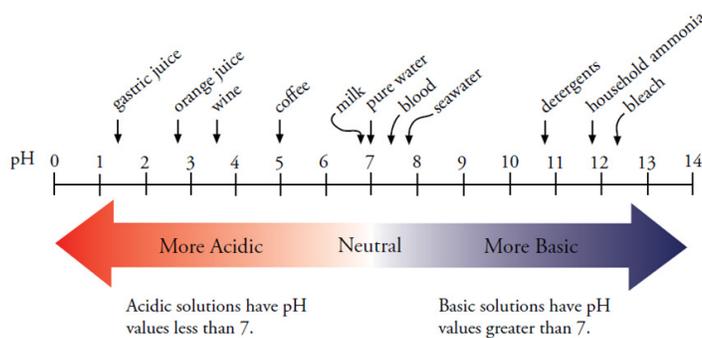
Carbon dioxide dissolved in water at a low concentration (0.2%–1.0%) creates carbonic acid (H_2CO_3), which causes the water to have a slightly sour taste with a pH between 3 and 4. An alkaline salt, such as sodium bicarbonate (NaHCO_3), may be added to soda water to reduce its acidity.

The amount of a gas like carbon dioxide that can be dissolved in water is described by Henry's Law. Water is chilled, optimally to just above freezing, in order to permit the maximum amount of carbon dioxide to dissolve in it. Higher gas pressure and lower temperature cause more gas to dissolve in the liquid. When the temperature is raised or the pressure is reduced (as happens when a container of carbonated water is opened,) carbon dioxide comes out of solution, in the form of bubbles.

Effervescence is the escape of gas from an aqueous solution and the foaming or fizzing that results from a release of the gas. The making of beer, wine, or champagne, by fermentation is thus also accompanied by effervescence of CO_2 from the barrel where the process occurs.

Effervescence can also be observed when opening a bottle of champagne, beer or carbonated beverages such as soft drinks. The visible bubbles are produced by the escape from solution of the dissolved gas.

This process is generally represented by the following reaction, where a pressurized dilute solution of carbonic acid in water releases gaseous carbon dioxide at decompression:



(8) Liming of Soil



Ammonia based fertilizer can lead to acidification of soil which could be adjusted by adding appropriate liming substances.

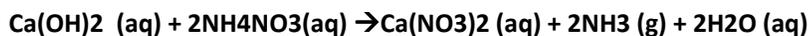
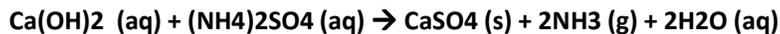


Macronutrients	Secondary Nutrients	Micronutrients
Nitrogen (N)	Calcium (Ca)	Zinc (Zn)
		Iron (Fe)
Phosphorus (P)	Magnesium (Mg)	Manganese (Mn)
		Copper (Cu)
Potassium (K)	Sulfur (S)	Boron (B)
		Molybdenum (Mo)

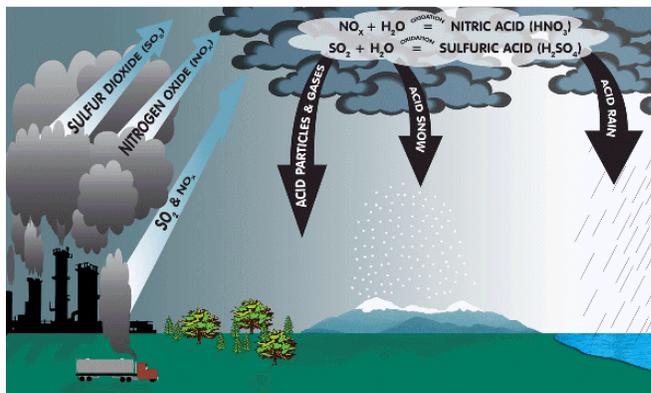
Common name	Chemical name	Chemical formula	cce
Ground limestone	Calcium carbonate	CaCO ₃	100
Dolomitic limestone	Calcium magnesium carbonate	CaMg(CO ₃) ₂	109
Hydrated lime	Calcium hydroxide	Ca(OH) ₂	135
Burned lime	Calcium	CaO	179

Limestone neutralizes any excess H⁺ in the soil whereas hydrated lime Ca(OH)₂ & quick lime CaO could potentially change the soil pH to alkaline. The primary purpose of liming is to adjust the pH and that of fertilizing is to supplement soil nutrient and they should not be

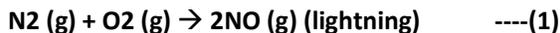
applied at the same time (especially for hydrated & quick lime) as ammonia salt react with alkaline which releases ammonia gas leading to loss of nutrient.



(9) Acid rain formation

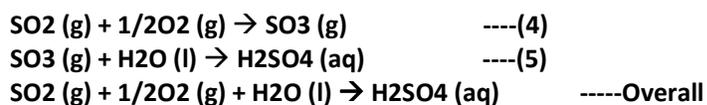


Nitric oxide (NO) is formed during lightning storms by the reaction of nitrogen and oxygen, two common atmospheric gases (Equation 1). In air, NO is oxidized to nitrogen dioxide (NO₂) (Equation 2), which in turn reacts with water to give nitric acid (HNO₃) (Equation 3). This acid dissociates in water to yield hydrogen ions and nitrate ions, lowering the pH of the solution.



About one-fourth of the acidity of rain is accounted for by nitric acid (HNO₃). What about the other 75% of the acidity of rain? Most is accounted for by the presence of sulfuric acid (H₂SO₄) in rainwater.

Although sulfuric acid may be produced naturally in small quantities from biological decay and volcanic activity, it is produced almost entirely by human activity, especially the combustion of sulfur-containing fossil fuels in power plants. When these fossil fuels are burned, the sulfur contained in them reacts with oxygen from the air to form sulfur dioxide (SO₂). Combustion of fossil fuels accounts for approximately 80% of the total atmospheric SO₂ in the United States. Sulfur dioxide, like the oxides of carbon and nitrogen, reacts with water to form sulfuric acid (Equation 4,5).



Sulfuric acid is a strong acid, so it readily dissociates in water, to give an H⁺ ion and an HSO₄⁻ ion. Unfortunately, human industrial activity produces additional acid-forming compounds in far greater quantities than the natural sources of acidity. In some areas of the United States, the pH of rainwater can be 3.0 or lower, approximately 1000 times more acidic than normal rainwater. When rainwater is too acidic, it can cause problems such as killing freshwater fish, damaging crops, and eroding buildings and monuments.

-----End of Summary-----