Unit 10

Inorganic Materials

This unit online, with its interactives:
science-playground.com/s/9a10

Turn on a tap in a city and the water sometimes smells faintly of a swimming pool: chlorine has been added to kill the germs in it. The soap you wash your hands with was made from sodium hydroxide, a chemical so strong it can burn skin. Your own stomach is making hydrochloric acid as you read this, and the salt in your dal is half of a poisonous gas, chlorine, joined to a metal that bursts into flame in water, sodium. Each of these substances has its own character: what it reacts with, what it makes, and why we can use it.

This unit looks at seven substances found in every laboratory and most homes, **oxygen**, **sulphur**, **chlorine**, **water**, **hydrochloric acid**, **sodium hydroxide and sodium chloride**, and at how each reacts, with balanced equations. On the way you will do two practicals on how fast reactions go. Then it turns to **alloys**: why your plate is steel and your gagri brass, and how a little of one metal changes another.

10.1 General chemical properties of oxygen, sulphur, chlorine, water, HCl, NaOH and NaCl

The **chemical properties** of a substance are the ways it reacts with other substances to form new ones: what it combines with, what it pushes out of a compound, whether it behaves as an acid or a base. To describe them clearly, chemists write **balanced equations**, with the conditions over the arrow (Δ for heat). Before you read on, two words you will meet often:

Oxidation is the addition of oxygen to a substance (or the removal of hydrogen from it). **Reduction** is the removal of oxygen (or the addition of hydrogen). A substance that gives oxygen to another, or takes hydrogen from it, is an **oxidizing agent**; one that takes oxygen away, or gives hydrogen, is a **reducing agent**. In $\mathrm{CuO} + \mathrm{H_2} \rightarrow \mathrm{Cu} + \mathrm{H_2O}$, copper oxide is reduced to copper and hydrogen is oxidized to water: copper oxide is the oxidizing agent, hydrogen the reducing agent.

Oxygen

Oxygen (O_2) is the most reactive gas in the air. It doesn't burn itself, but it **supports burning**, and it combines with almost every other element to form **oxides**. It is a strong **oxidizing agent**.

1. With metals: basic oxides. Most metals burn in oxygen, or slowly tarnish in it, forming metal oxides:

$$2Mg + O_2 \xrightarrow{\Delta} 2MgO \qquad 2Ca + O_2 \xrightarrow{\Delta} 2CaO$$

$$3Fe + 2O_2 \xrightarrow{\Delta} Fe_3O_4 \text{ (iron wool burning in oxygen, with sparks)}$$

$$4Al + 3O_2 \rightarrow 2Al_2O_3 \text{ (a thin film that protects aluminium pots)}$$

Iron left in moist air rusts instead: it combines slowly with oxygen **and water** to form hydrated iron(III) oxide, $Fe_2O_3 \cdot xH_2O$, the brown rust.

2. With non-metals: acidic oxides.

$$C + O_2 \rightarrow CO_2 \qquad S + O_2 \rightarrow SO_2 \qquad 4P + 5O_2 \rightarrow 2P_2O_5$$

$$2H_2 + O_2 \rightarrow 2H_2O \qquad N_2 + O_2 \xrightarrow{\text{very high temperature}} 2NO$$

Nitrogen is so unreactive that it joins oxygen only in a lightning flash or a hot car engine.

> CAREFUL
>
> **Phosphorus pentoxide: P_2O_5 or P_4O_{10}?**
>
> White phosphorus is made of P_4 molecules, and the oxide of phosphorus is made of P_4O_{10} molecules, a cage of four phosphorus atoms joined by oxygen: $P_4 + 5O_2 \rightarrow P_4O_{10}$. P_2O_5 is its simplest formula, half of the molecule. Exams write P_2O_5 and $4P + 5O_2 \rightarrow 2P_2O_5$, and that equation earns the marks; the two say the same thing, since $2P_2O_5$ has the same atoms as P_4O_{10}.

3. With compounds. Oxygen oxidizes many compounds, burning them to oxides:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O \qquad 2CO + O_2 \rightarrow 2CO_2$$

$$2H_2S + 3O_2 \rightarrow 2SO_2 + 2H_2O \qquad 4NH_3 + 3O_2 \rightarrow 2N_2 + 6H_2O$$

4. Into ozone. In a spark or in the Sun's ultraviolet light high in the air, oxygen molecules join into **ozone**, O_3: $3O_2 \rightarrow 2O_3$.

Four kinds of oxide

What an oxide does in water, and with acids and bases, sorts it into one of four kinds:

5. Numerical problems

a. A gold necklace of 18 carats has a mass of 35 g. Find the mass of pure gold in it.

b. A craftsman in Lalitpur melts 3.9 kg of copper with 1.1 kg of tin to cast a bell. What is the percentage of copper in the alloy? Is it ordinary bronze or bell metal?

c. **Think:** Riya adds 0.5 g of marble powder to 40 mL of dilute hydrochloric acid, and Nabin adds 0.5 g of marble chips to another 40 mL of the same acid. Riya's fizzing stops within a minute; Nabin's goes on for many minutes. Nabin says, "My chips must give more carbon dioxide, since they fizz for longer." Is he right? Explain.

Found a mistake? Please tell us.

Kind	What it does	Examples
Basic oxide (oxides of metals)	reacts with acids to give a salt and water; if it dissolves, it gives an alkali, which turns red litmus blue	Na_2O, MgO, CaO, CuO
Acidic oxide (oxides of most non-metals)	reacts with bases to give a salt and water; in water it gives an acid, which turns blue litmus red	CO_2, SO_2, SO_3, P_2O_5
Amphoteric oxide	reacts with both acids and bases	ZnO, Al_2O_3
Neutral oxide	neither acidic nor basic	CO, NO, N_2O, H_2O

$$CaO + H_2O \rightarrow Ca(OH)_2 \qquad SO_2 + H_2O \rightarrow H_2SO_3$$

$$ZnO + 2HCl \rightarrow ZnCl_2 + H_2O \qquad ZnO + 2NaOH \rightarrow Na_2ZnO_2 + H_2O$$

Check 1

An element burns in oxygen, and its oxide dissolves in water to give a solution that turns red litmus blue. The element is most likely:

a sulphur b carbon
c calcium d phosphorus

Making oxygen with a catalyst

In the laboratory, oxygen is made by letting hydrogen peroxide break down into water and oxygen. On its own, hydrogen peroxide decomposes so slowly that a bottle keeps for months. Add a pinch of black **manganese dioxide**, and it froths with oxygen at once:

$$2H_2O_2 \xrightarrow{MnO_2} 2H_2O + O_2\uparrow$$

The manganese dioxide is written over the arrow, not among the reactants, because it is not used up. It is a **catalyst**: a substance that changes the rate of a chemical reaction without itself being permanently changed. The **rate of a reaction** is how fast the reactants turn into products: the amount of product formed (here, the volume of oxygen) in each second or minute.

Why does anything change the rate? Particles can react only when they **collide**, and only a collision with enough energy, the **activation energy**, breaks the old bonds so the new ones can form. A catalyst gives the reaction an easier path with a lower activation energy, so many more collisions succeed. Anything that makes the particles collide **more often** or **harder** also speeds a reaction up:

Exercise

1. Choose the correct option

a. Which of these is an amphoteric oxide?

- **a** CaO
- **b** SO_2
- **c** ZnO
- **d** CO

b. When iron reacts with dilute hydrochloric acid, the products are:

- **a** $FeCl_3$ and H_2
- **b** $FeCl_2$ and H_2
- **c** $FeCl_2$ and Cl_2
- **d** no reaction

c. Which property makes nichrome suitable for the heating coil of an electric heater?

- **a** low resistance and low melting point
- **b** it is magnetic
- **c** high resistance, and it doesn't oxidize when red hot
- **d** it is a good conductor of heat only

2. Give reasons

a. Dry chlorine doesn't bleach a dry flower, but moist chlorine does.

b. Sulphur powder is sprinkled over mercury spilled from a broken thermometer.

c. Alloys are used more than pure metals in daily life.

3. Differentiate between

i. Acidic oxide and basic oxide

ii. Substitutional alloy and interstitial alloy

4. Answer the following questions

a. Write the balanced equations for the reactions of chlorine with (i) hydrogen, (ii) iron, (iii) cold dilute sodium hydroxide, (iv) slaked lime, (v) potassium iodide.

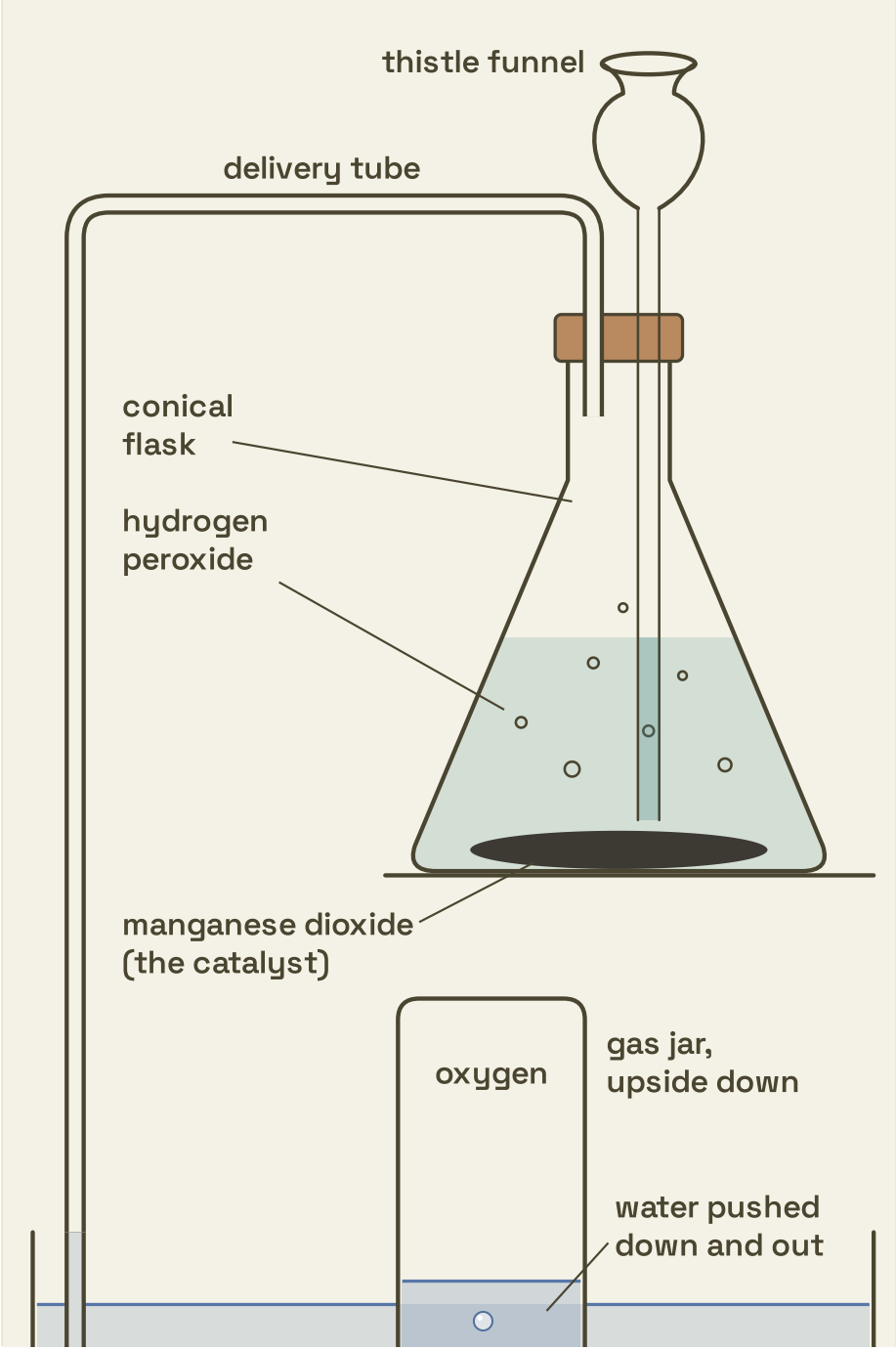
b. How does dilute hydrochloric acid react with (i) zinc, (ii) sodium carbonate, (iii) copper oxide, (iv) silver nitrate? Write balanced equations and what you would see.

c. Complete and balance, and name each product: (i) $Zn + S \rightarrow$, (ii) $Na + H_2O \rightarrow$, (iii) $CaO + H_2O \rightarrow$, (iv) $NaCl + H_2SO_4$ (conc., warm) $\rightarrow$, (v) $NH_4Cl + NaOH$ (warm) $\rightarrow$, (vi) $NaCl + H_2O$ (brine, electrolyzed) $\rightarrow$.

d. What is an alloy? Write the composition and one use each of brass, bronze, stainless steel, solder and duralumin. What is an amalgam? Give one example.

IN SHORT

- **Oxygen** supports burning and oxidizes most elements: metals give **basic oxides**, non-metals **acidic oxides**; ZnO and Al_2O_3 are **amphoteric**, CO and NO **neutral**.
- **Catalyst**: changes the rate of a reaction and is left unchanged (MnO_2 in $2H_2O_2 \rightarrow 2H_2O + O_2$). Heat, higher concentration and more surface area also speed reactions up, by making successful collisions more frequent.
- **Sulphur** burns to SO_2 with a blue flame and forms sulphides with metals (FeS, ZnS, HgS), H_2S with hydrogen and CS_2 with carbon; hot concentrated acids oxidize it.
- **Chlorine** forms chlorides with hydrogen and metals; with water it gives HCl and HOCl, whose nascent oxygen **bleaches** (moist only); it gives bleach with NaOH, bleaching powder with slaked lime, and displaces bromine and iodine.
- **Water** gives alkalis and hydrogen with reactive metals, alkalis and acids with oxides, hydrates with salts (white $CuSO_4$ turns blue), and H_2 and O_2 by electrolysis.
- **Hydrochloric acid**, a strong acid, gives hydrogen with metals above hydrogen, salt and water with bases, CO_2 with carbonates, white AgCl with silver nitrate, white fumes with ammonia.
- **Sodium hydroxide**, a strong alkali, neutralizes acids, absorbs CO_2, precipitates colored metal hydroxides, frees ammonia from ammonium salts, dissolves zinc and aluminium, and makes soap from fats.
- **Sodium chloride**, a neutral salt, gives HCl with concentrated H_2SO_4 and AgCl with $AgNO_3$, and is the raw material for NaOH, Cl_2, H_2, baking soda and washing soda.
- An **alloy** is a homogeneous mixture of a metal with other metals or a non-metal: **substitutional** (brass) or **interstitial** (steel); **ferrous**, **non-ferrous** or **amalgams**. Each is chosen for a property its pure metal lacks.



- **Heat.** Warmer particles move faster, so they collide more often, and many more of the collisions have enough energy. For many reactions a rise of 10 °C roughly doubles the rate.
- **Concentration.** A stronger solution has more reacting particles in each milliliter, so they collide more often.
- **Surface area.** Only the particles on the surface of a solid can be hit. A powder has far more surface than a lump of the same mass, so it reacts faster.

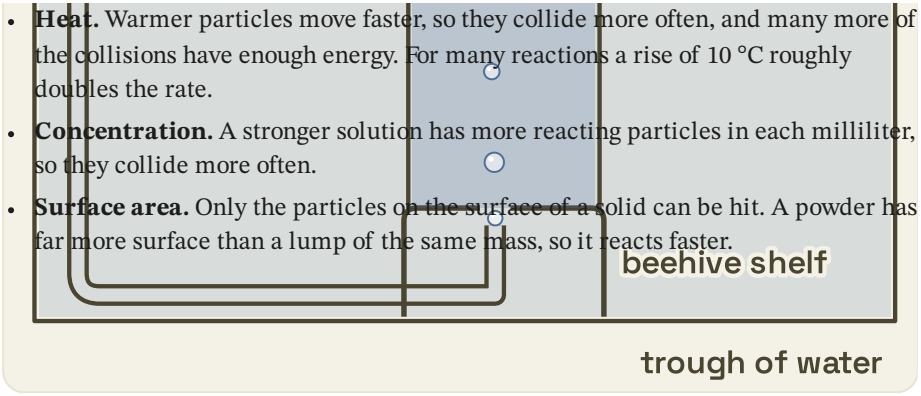


Figure 10.1 *Making oxygen without heating: hydrogen peroxide is poured through the thistle funnel onto manganese dioxide, and the oxygen is collected over water.*

Sulphur

Sulphur is a yellow, brittle non-metal, found near volcanoes and hot springs and taken out of crude oil and natural gas. Its molecules are rings of eight atoms, S_8, but equations usually write it simply as S. Sulphur is less reactive than oxygen, but when it is heated it combines with most metals and many non-metals.

1. With oxygen. Heated, sulphur melts to an amber liquid, which darkens to red-brown as it gets hotter, and burns with a blue flame, faint in air and bright in oxygen, giving sulphur dioxide, a choking gas:

$$S + O_2 \rightarrow SO_2$$

2. With metals: sulphides. Heated with iron, zinc or copper, sulphur forms the metal's sulphide; mercury and sulphur react even cold:

$$Fe + S \xrightarrow{\Delta} FeS \qquad Zn + S \xrightarrow{\Delta} ZnS \qquad 2Cu + S \xrightarrow{\Delta} Cu_2S \qquad Hg + S \rightarrow HgS$$

That last reaction is why sulphur powder is sprinkled over spilled mercury: it turns the poisonous droplets into solid mercury sulphide, which doesn't give off vapor.

3. With hydrogen and carbon. Hydrogen passed over boiling sulphur gives hydrogen sulphide, the gas with the smell of rotten eggs; red-hot charcoal gives carbon disulphide:

$$H_2 + S \xrightarrow{\Delta} H_2S \qquad C + 2S \xrightarrow{\text{red hot}} CS_2$$

4. With concentrated acids. Hot concentrated sulphuric acid and nitric acid, both strong oxidizing agents, oxidize sulphur:

$$S + 2H_2SO_4 \xrightarrow{\Delta} 3SO_2 + 2H_2O \qquad S + 6HNO_3 \xrightarrow{\Delta} H_2SO_4 + 6NO_2 + 2H_2O$$

Dilute acids have no effect on sulphur.

Alloy	Made of	Property that matters	Used for
duralumin	Al + Cu, Mg, Mn	light and strong	aircraft, bicycle and car parts
solder	Sn + Pb (or Sn + Ag + Cu)	low melting point	joining wires and pipes
nichrome	Ni + Cr	high resistance, no oxidizing when hot	heating coils
invar	Fe + Ni	hardly expands when heated	clocks, measuring tapes
dental amalgam	Hg + Ag, Sn, Cu	soft when mixed, sets hard	filling teeth

WORKED EXAMPLE

How much gold in a bangle?

A bangle of 22-carat gold has a mass of 23.3 g (about two tolas). How much pure gold does it contain, and what percentage is gold?

Answer

Carats count parts of gold in 24: 22 carats means 22 parts in 24.

$$\text{Percentage of gold} = \frac{22}{24} \times 100\% = 91.7\%$$

$$\text{Mass of gold} = \frac{22}{24} \times 23.3\ \text{g} = 21.4\ \text{g}$$

21.4 g of gold, 91.7%. The other 1.9 g is copper and silver, which make it hard enough to wear every day.

WORKED EXAMPLE

Making brass

A brass maker in Patan wants 6 kg of brass with 65% copper and 35% zinc. How much of each metal must be melted together?

Answer

$$\text{Copper} = \frac{65}{100} \times 6\ \text{kg} = 3.9\ \text{kg} \qquad \text{Zinc} = \frac{35}{100} \times 6\ \text{kg} = 2.1\ \text{kg}$$

3.9 kg of copper and 2.1 kg of zinc (check: $3.9 + 2.1 = 6.0$ kg).

Application



Figure 10.6 *An X-ray of a tooth with a dental amalgam filling: the metal stops X-rays, so it shows bright white. About half of an amalgam filling is mercury, held firmly in the alloy; still, under the Minamata Convention on mercury, many countries are now cutting down its use.*

Photo: Bernard bill5, CC BY-SA 3.0

Check 5

Which pair is an interstitial alloy and a substitutional alloy, in that order?

a brass and steel
b steel and brass
c bronze and brass
d solder and bronze

Alloy	Made of	Property that matters	Used for
steel	Fe + up to about 2% C	hard, strong, cheap	buildings, bridges, vehicles, tools
stainless steel	Fe + Cr (+ Ni)	doesn't rust	utensils, pressure cookers, surgical tools
brass	Cu + Zn	golden, easy to cast and machine, doesn't rust	gagri, puja items, taps, screws
bronze, bell metal	Cu + Sn	hard, rings when struck	kansa plates, bells, statues, medals
German silver	Cu + Ni + Zn	silvery, hard	cutlery, ornaments, musical instruments

Uses. Most sulphur is turned into sulphuric acid, for fertilizers. It also hardens rubber for tires (vulcanization), and goes into matches, gunpowder, fungicides sprayed on fruit and vegetables, and some skin ointments.



Figure 10.2 *The crater of the Ijen volcano in Indonesia. Sulphur crystallizes as yellow crusts around the vents where volcanic gas escapes, and miners (the small figures on the left) break it off and carry it out by hand.*

Photo: Sémhur, CC BY-SA 4.0

Chlorine

Chlorine (Cl_2) is a greenish-yellow gas with a sharp, choking smell, about two and a half times as dense as air, and poisonous. Its name comes from the Greek *chloros*, pale green. It is one of the most reactive non-metals: its atoms need only one more electron to fill their outer shell. In the laboratory it is made by heating manganese dioxide with concentrated hydrochloric acid:

$$MnO_2 + 4HCl \xrightarrow{\Delta} MnCl_2 + Cl_2 \uparrow + 2H_2O$$

1. With hydrogen. A mixture of hydrogen and chlorine reacts slowly in dim light and **explodes** in bright sunlight, giving hydrogen chloride:

$$H_2 + Cl_2 \xrightarrow{\text{sunlight}} 2HCl$$

2. With metals: chlorides. Hot sodium burns in chlorine with a yellow flame, and hot iron wool and copper glow in it:

$$2Na + Cl_2 \rightarrow 2NaCl \qquad 2Fe + 3Cl_2 \rightarrow 2FeCl_3 \qquad Cu + Cl_2 \rightarrow CuCl_2$$

Chlorine gives iron its higher valency, 3, forming iron(III) chloride.

3. With water, and how it bleaches. Chlorine dissolves in water and partly reacts with it, giving hydrochloric acid and **hypochlorous acid**, HOCl:

$$Cl_2 + H_2O \rightleftharpoons HCl + HOCl$$

Hypochlorous acid is unstable and gives up its oxygen easily, written as **nascent oxygen**, [O], a very reactive form:

$$HOCl \rightarrow HCl + [O]$$

The nascent oxygen oxidizes colored substances into colorless ones: colored matter + [O] → colorless matter. That is **bleaching by oxidation**, and the same oxidizing kills bacteria. Because it needs water, **dry chlorine doesn't bleach**: a dry colored flower

or cloth stays colored in dry chlorine and loses its color in moist chlorine. In sunlight, chlorine water slowly gives off oxygen: $2HOCl \rightarrow 2HCl + O_2$.

4. With alkalis. With cold, dilute sodium hydroxide, chlorine gives sodium hypochlorite, the bleach sold in bottles; with hot, concentrated sodium hydroxide it gives sodium chlorate. With slaked lime it gives **bleaching powder**:

$$Cl_2 + 2NaOH \xrightarrow{\text{cold, dilute}} NaCl + NaOCl + H_2O$$

$$3Cl_2 + 6NaOH \xrightarrow{\text{hot, conc.}} 5NaCl + NaClO_3 + 3H_2O$$

$$Ca(OH)_2 + Cl_2 \rightarrow CaOCl_2 + H_2O$$

CAREFUL

The formula of bleaching powder

Exams write bleaching powder as $CaOCl_2$ (calcium oxychloride), and that equation earns the marks. Chemists today know it is really a mixture, mainly calcium hypochlorite, $Ca(OCl)_2$, with calcium chloride and slaked lime: $2Ca(OH)_2 + 2Cl_2 \rightarrow Ca(OCl)_2 + CaCl_2 + 2H_2O$. Both equations are balanced; the hypochlorite part is what bleaches and disinfects.

5. Pushing out bromine and iodine. Chlorine is more reactive than bromine and iodine, so it displaces them from their salts. Chlorine water added to colorless potassium iodide solution turns it brown with iodine:

$$Cl_2 + 2KBr \rightarrow 2KCl + Br_2 \qquad Cl_2 + 2KI \rightarrow 2KCl + I_2$$

6. With hydrogen sulphide. Chlorine takes hydrogen away from hydrogen sulphide (it is an oxidizing agent), leaving yellow sulphur: $H_2S + Cl_2 \rightarrow 2HCl + S$.

Uses. Killing germs in drinking water and swimming pools; bleaching paper and cotton; making bleaching powder, PVC plastic, hydrochloric acid and many medicines and pesticides.

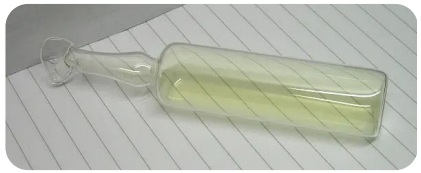


Figure 10.3 *Chlorine sealed in a glass tube. Its pale greenish-yellow color gave it its name.*
Photo: W. Oelen, CC BY-SA 3.0

- **Bronze**: copper and tin. **Bell metal**, a bronze with more tin (about 78% copper, 22% tin), rings well and is used for bells and kansa plates.
- **Gunmetal**: copper with about 8–10% tin and 2–4% zinc, for valves, gears and statues.
- **German silver** (nickel silver): about 60% copper, 20% nickel and 20% zinc; silvery, with no silver in it.

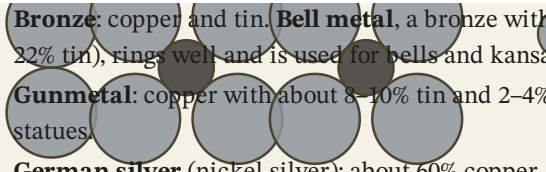
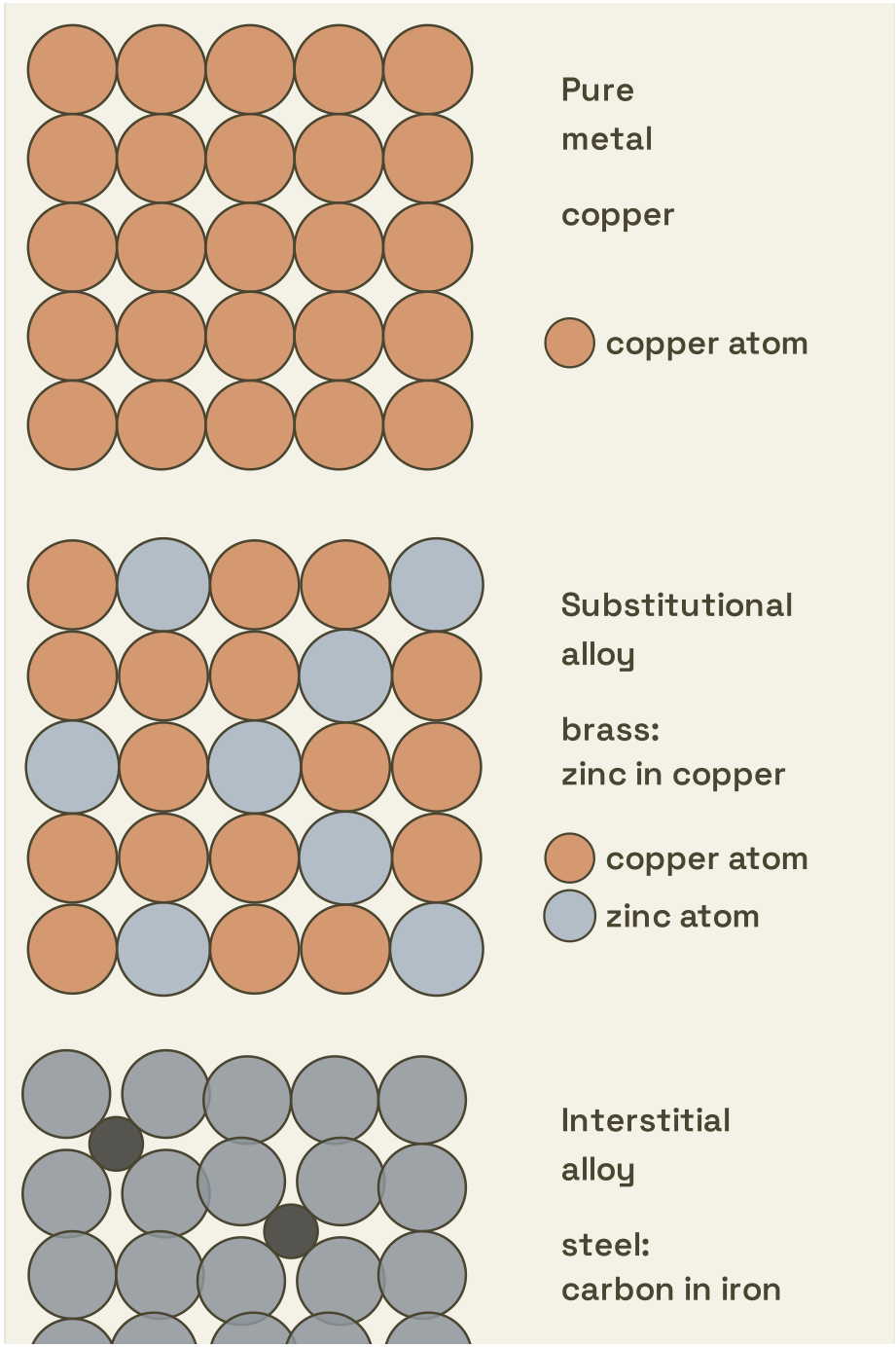


Figure 10.5 *A slice through a crystal of each, with the atoms' sizes to scale (radii: copper 128 pm, zinc 134 pm, iron 126 pm, carbon 77 pm). In brass, zinc atoms take some of copper's places; in steel, the small carbon atoms squeeze into the gaps between iron atoms, pushing their neighbors apart a little. Either way the rows are no longer perfectly even, so they slide less easily.*

- **Duralumin**: aluminium with about 4% copper and a little magnesium and manganese; light and strong.
- **Magnalium**: aluminium and magnesium; light, for aircraft parts and balance beams.
- **Solder**: tin and lead (63% tin, 37% lead is common), melting at about 183 °C, or lead-free tin with a little silver and copper; joins wires and pipes.
- **Nichrome**: 80% nickel and 20% chromium; high resistance, and doesn't oxidize when red hot, so it is the heating wire of heaters, irons and kettles.
- **Gold of 22 carats**: 22 parts in 24 gold, the rest copper and silver.

3. Amalgams are alloys of **mercury** with other metals. **Dental amalgam**, mercury mixed with silver, tin and copper, is soft enough to press into a tooth and then sets hard. **Sodium amalgam** is used in chemistry as a gentle reducing agent. Iron and platinum don't form amalgams, which is why mercury is stored in iron flasks.



A STORY FROM SCIENCE

From a choking gas to clean water

The Swedish chemist Carl Wilhelm Scheele first made chlorine in 1774, by heating the mineral pyrolusite (manganese dioxide) with hydrochloric acid, the same reaction as above. For years chemists thought the gas was a compound with oxygen in it; in 1810 Humphry Davy showed it was an element and named it. In 1847 the doctor Ignaz Semmelweis, in Vienna, had doctors wash their hands in a solution of chlorinated lime (bleaching powder) before examining mothers, and deaths from fever after childbirth in his ward fell from about 18 in 100 to about 2. In 1908 Jersey City became the first city in the United States to add chlorine to all its drinking water, every day, and within ten years nearly every large American city did the same. Few chemicals have saved more lives. The same gas was also the first poison gas to kill on a large scale in war, at Ypres in Belgium, on 22 April 1915.

Check 2

A piece of dry colored cloth is put into a jar of dry chlorine, and another piece, wet, into a second jar. What happens?

a both lose their color
b neither loses its color
c only the wet cloth loses its color
d only the dry cloth loses its color

Water

Water (H_2O) is a neutral compound: pure water doesn't change the color of litmus. Yet it reacts with many substances, and it is the solvent in which most reactions on Earth happen.

1. With metals. Very reactive metals react with cold water, giving an alkali and hydrogen; less reactive ones need hot water or steam:

$$2Na + 2H_2O \rightarrow 2NaOH + H_2\uparrow \qquad Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2\uparrow$$

$$Mg + H_2O\,(steam) \rightarrow MgO + H_2 \qquad 3Fe + 4H_2O\,(steam) \rightleftharpoons Fe_3O_4 + 4H_2$$

Copper, silver and gold don't react with water at all, which is why they make good water pipes and long-lasting ornaments.

2. With non-metals. Steam passed over red-hot coke gives **water gas**, a fuel (and chlorine dissolves in water to give chlorine water, as above):

$$C + H_2O \xrightarrow{1{,}000\ ^\circ C} CO + H_2$$

3. With oxides. Water turns basic oxides into alkalis and acidic oxides into acids. Water poured on quicklime hisses and steams as it forms slaked lime, the white of whitewash:

$$CaO + H_2O \rightarrow Ca(OH)_2 + \text{heat} \qquad Na_2O + H_2O \rightarrow 2NaOH$$

$$CO_2 + H_2O \rightleftharpoons H_2CO_3 \qquad SO_3 + H_2O \rightarrow H_2SO_4$$

4. With salts: water of crystallization. White anhydrous copper sulphate takes up water and turns blue. This color change is the usual test for water:

$$CuSO_4\,(\text{white}) + 5H_2O \rightarrow CuSO_4 \cdot 5H_2O\,(\text{blue})$$

5. By electricity. Water with a little acid in it is split by electrolysis into hydrogen and oxygen, two volumes to one: $2H_2O \rightarrow 2H_2 + O_2$ (you did this in Unit 8).

Hydrochloric acid (HCl)

Hydrogen chloride is a colorless gas with a sharp smell, made of HCl molecules. It dissolves very readily in water, and the solution is **hydrochloric acid**. In water the molecules split into ions, $HCl \rightarrow H^+ + Cl^-$, almost completely, so hydrochloric acid is a **strong acid**. Concentrated hydrochloric acid is about 36% HCl and gives off choking fumes; the dilute acid used in most experiments is much weaker. Your stomach makes dilute hydrochloric acid to digest food and kill germs: gastric juice has a pH of about 1 to 3.

> CAREFUL
>
> **Hydrogen chloride or hydrochloric acid?**
>
> Dry hydrogen chloride gas doesn't turn dry blue litmus red and doesn't conduct electricity: its molecules aren't split into ions. Only in water does it form the H^+ ions that make it an acid. So write HCl(g) for the gas and HCl(aq) for the acid when the difference matters; the properties below are those of the acid.

1. Acidic nature. It turns blue litmus red and methyl orange red.

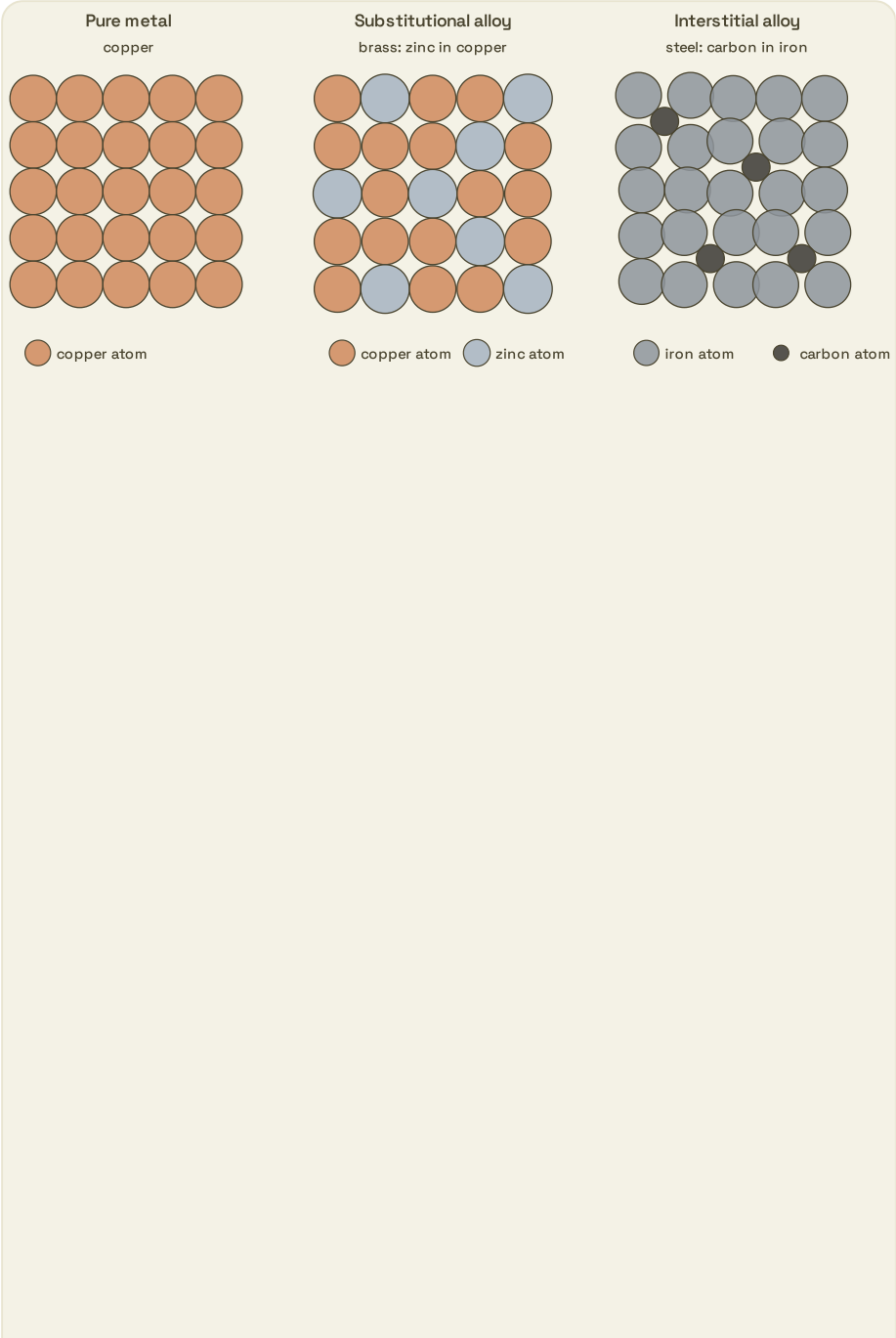
2. With metals above hydrogen in the reactivity series, it gives a salt and hydrogen:

$Zn + 2HCl \rightarrow ZnCl_2 + H_2\uparrow$ $\quad Mg + 2HCl \rightarrow MgCl_2 + H_2\uparrow$ $\quad Fe + 2HCl \rightarrow Fe$

Copper, silver and gold, below hydrogen, don't react with it.

3. With bases and metal oxides (neutralization), it gives a salt and water:

$NaOH + HCl \rightarrow NaCl + H_2O$ $\quad CuO + 2HCl \rightarrow CuCl_2 + H_2O$ $\quad CaO + 2HCl \rightarrow$



10.2 Alloys

Introduction and types

An **alloy** is a homogeneous mixture of a metal with one or more other metals, or with a non-metal such as carbon, made by melting them together and letting them cool. Pure metals are often too soft, rust too easily or melt at the wrong temperature for the job; mixing in a little of something else changes that. An alloy is usually **harder and stronger** than its main metal, often **resists corrosion** better, often **melts at a lower temperature**, and may have a different color, conductivity or magnetism.

Why harder? In a pure metal all the atoms are the same size and sit in regular rows, and the rows can slide over each other when the metal is hammered or bent. In an alloy, atoms of a different size upset the rows, so they can't slide so easily. How the second element fits in gives the two kinds of alloy shown in Figure 10.5:

- In a **substitutional alloy**, atoms of the added metal, about the same size as the main metal's, **take the places** of some of its atoms. Brass is one: zinc atoms take the places of some copper atoms.
- In an **interstitial alloy**, much **smaller atoms**, such as carbon, fit into the **gaps** between the main metal's atoms. Steel is one: carbon atoms sit in the gaps between iron atoms.

Alloys are also grouped by their main metal:

1. Ferrous alloys contain iron as the main metal.

- **Steel**: iron with a little carbon (up to about 2%). More carbon makes it harder but more brittle: mild steel for rods and sheets, high-carbon steel for knives, springs and tools.
- **Stainless steel**: iron with at least 10.5% chromium, often with nickel; the common "18/8" steel of plates and pressure cookers has 18% chromium and 8% nickel. The chromium forms an invisible oxide film that heals itself when scratched.
- **Invar**: iron with 36% nickel, which hardly expands when heated; used in clocks, measuring tapes and precision instruments.
- **Alnico**: iron with aluminium, nickel and cobalt, for strong permanent magnets.

2. Non-ferrous alloys contain no iron (or very little) as the main metal.

- **Brass**: copper and zinc, usually about two parts copper to one of zinc.

CAREFUL

Iron gives iron(II) chloride

With hydrochloric acid, iron forms **iron(II) chloride**, $FeCl_2$ (a pale green solution), not $FeCl_3$. The acid's hydrogen ions are not a strong enough oxidizing agent to take iron to valency 3. Chlorine gas, a much stronger oxidizing agent, does give $FeCl_3$, as you saw above.

4. With carbonates and hydrogen carbonates, it fizzes, giving a salt, water and carbon dioxide (which turns lime water milky):

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2 \uparrow$$

$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2 \uparrow \qquad NaHCO_3 + HCl \rightarrow NaCl + H_2O + C$$

An antacid tablet works this way: a carbonate or hydrogen carbonate neutralizes some of the stomach's acid.

5. With silver nitrate: the test for chloride. It gives a white precipitate of silver chloride, which darkens in sunlight:

$$AgNO_3 + HCl \rightarrow AgCl \downarrow + HNO_3$$

6. With ammonia. Hydrogen chloride gas and ammonia gas meet in dense white fumes of ammonium chloride (the white ring in the Graham's law tube of Unit 7):

$$NH_3 + HCl \rightarrow NH_4Cl$$

Uses. Cleaning rust and scale off steel before it is coated (pickling); making chlorides and PVC; cleaning tiles and toilets (as a dilute acid); in laboratories everywhere. Mixed with nitric acid (3 : 1) it makes **aqua regia**, which dissolves gold.

Check 3

Which of these does NOT react with dilute hydrochloric acid?

a zinc
b marble
c copper
d sodium hydroxide

Sodium hydroxide (NaOH)

Sodium hydroxide, **caustic soda**, is a white solid sold as pellets or flakes. It dissolves in water with a lot of heat, and the solution feels soapy, because it turns the fats of your skin into soap. It is **caustic**: it burns skin and is very dangerous to the eyes. In water it is fully split into Na^+ and OH^- ions, so it is a **strong alkali**: it turns red litmus blue and phenolphthalein pink. It is made by the electrolysis of brine 🕮 Grade 9.

1. With acids (neutralization), it gives a salt and water:

$$NaOH + HCl \rightarrow NaCl + H_2O \qquad 2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$$

2. With acidic oxides. It absorbs carbon dioxide and sulphur dioxide. A bottle of solid caustic soda left open turns damp (it takes water from the air: it is **deliquescent**) and crusty with sodium carbonate:

$$2NaOH + CO_2 \rightarrow Na_2CO_3 + H_2O \qquad 2NaOH + SO_2 \rightarrow Na_2SO_3 + H_2O$$

3. With salts of heavy metals: colored precipitates. Most metal hydroxides don't dissolve in water, so sodium hydroxide solution added to a metal's salt solution throws down its hydroxide, each with its own color, which helps to identify the metal:

$$CuSO_4 + 2NaOH \rightarrow Cu(OH)_2 \downarrow \text{ (blue)} + Na_2SO_4$$

$$FeSO_4 + 2NaOH \rightarrow Fe(OH)_2 \downarrow \text{ (dirty green)} + Na_2SO_4 \qquad FeCl_3 + 3NaOH \rightarrow Fe(OI$$

4. With ammonium salts: the test for ammonium. Warmed with an ammonium salt, it gives off ammonia, which turns damp red litmus blue:

$$NH_4Cl + NaOH \xrightarrow{\Delta} NaCl + NH_3 \uparrow + H_2O$$

5. With zinc and aluminium. Unlike most metals, zinc and aluminium dissolve in hot sodium hydroxide solution, giving hydrogen:

$$Zn + 2NaOH \rightarrow Na_2ZnO_2 + H_2 \uparrow \qquad 2Al + 2NaOH + 2H_2O \rightarrow 2NaAlO_2 + 3H_2 \uparrow$$

That is why an aluminium pot should never be cleaned with caustic soda or washing soda: the alkali eats into it.

CAREFUL

Sodium aluminate

Exams write the product as sodium aluminate, $NaAlO_2$, as above. In water it is really present as $Na[Al(OH)_4]$, and chemists write $2Al + 2NaOH + 6H_2O \rightarrow 2Na[Al(OH)_4] + 3H_2$. The two say the same thing: the second only counts the water joined to the aluminate.

6. With fats: soap. Boiled with oils or fats, sodium hydroxide makes soap and glycerol. This is **saponification**: fat + sodium hydroxide → soap + glycerol.

Uses. Making soap, detergents, paper and rayon; refining vegetable oil; clearing blocked drains; and in many laboratory tests.

Sodium chloride (NaCl)

Sodium chloride, common salt, is the **neutral salt** of a strong acid and a strong base: $NaOH + HCl \rightarrow NaCl + H_2O$. It is an ionic solid of Na^+ and Cl^- ions; its solution doesn't change litmus and conducts electricity. It comes from seawater, salt lakes and rock salt mines. Most table salt sold in Nepal is **iodized**: a tiny amount of an iodine compound is added, because too little iodine in food causes goiter and holds back children's growth and learning.



Figure 10.4 *Sodium hydroxide pellets. Left in the open, they soon take up water from the air and turn sticky.*
Photo: Walkerma, Public domain

1. With concentrated sulphuric acid. It gives hydrogen chloride gas, which is how hydrogen chloride is made in the laboratory:

$$NaCl + H_2SO_4 \xrightarrow{\text{warm}} NaHSO_4 + HCl \uparrow$$

Check 4

A few drops of sodium hydroxide solution are added to a solution, and a blue precipitate forms. The solution contained:

a iron(III) chloride
b copper sulphate
c sodium chloride
d iron(II) sulphate

2. With silver nitrate. A white precipitate of silver chloride, the test for chloride ions:

$$NaCl + AgNO_3 \rightarrow AgCl \downarrow + NaNO_3$$

3. By electrolysis. Molten salt gives sodium metal and chlorine; brine (a strong solution in water) gives sodium hydroxide, hydrogen and chlorine, the three chemicals of the chlor-alkali industry:

$$2NaCl\ (\text{molten}) \xrightarrow{\text{electricity}} 2Na + Cl_2 \qquad 2NaCl + 2H_2O \xrightarrow{\text{electricity}} 2NaOH + H_2$$

4. Into baking soda and washing soda (the Solvay process). Brine saturated with ammonia and carbon dioxide gives sodium hydrogen carbonate, baking soda, which on heating gives sodium carbonate, washing soda:

$$NaCl + NH_3 + CO_2 + H_2O \rightarrow NaHCO_3 + NH_4Cl \qquad 2NaHCO_3 \xrightarrow{\Delta} Na_2CO_3 + H_2O$$

Uses. In food, and to preserve pickles, meat and fish; in oral rehydration solution for diarrhea, with glucose and other salts; as the raw material for sodium hydroxide, chlorine, hydrochloric acid, sodium metal, baking soda and washing soda; to melt ice on roads in cold countries.