

**Layers of the Atmosphere**

The atmosphere is a gaseous layer that extends approximately 600km above the Earth’s surface.

The temperature changes mark the boundaries between each layer of the atmosphere.



* Troposphere 0-15km
  + Air pressure highest
  + Contains 75% by mass of the atmosphere
  + Gases are well mixed due to convection currents
  + Temperature decreases with altitude (15°C at surface to -50°C at boundary)
  + Boundary is called the tropopause
  + Rate of transfer of gases across the tropopause is slow
  + Tropopause is higher at the equator than the poles due to expansion of hot air below
  + Water vapour at the tropopause crystalises and is not lost between the layers
* Stratosphere 15-50km
  + Contains the ozone layer (from 10km to boundary)
  + Upper ozone absorbs low wavelength radiation (UV-B and UV-C)
  + Air pressure decreases with altitude (10kPa to 0.1kPa at boundary)
  + Uniform temperature for first 9km, temperature increases after this
  + Gases mix very little as convection currents are prevented
  + Air is dry and stable
  + Pollutants that enter usually remain for long periods
  + Boundary is called the stratopause
* Mesosphere 50-85km
  + Air pressure decreases
  + Temperature decreases
  + Few molecules absorb radiation so the zone is quite cold
  + Boundary is called the mesopause
* Thermosphere 85-600km
  + Temperature rises due to high frequency radiation
* Ionosphere
  + Region where ions are produced
  + Includes part of the mesosphere and thermosphere
  + Air is very thin (meaning large spaces between molecules)
  + Temperature reaches 1700°C
  + No distinct end to the atmosphere, fades gradually into the low pressures of space

**Composition of the Atmosphere**

The concentration of gas particles will drop with altitude, however the proportion of gases remains constant throughout the layers.

The atmosphere contains the following gases:

|  |  |  |  |
| --- | --- | --- | --- |
| **Gas** | **Concentration %(v/v)** | **Gas** | **Concentration %(v/v)** |
| Nitrogen | 78.09 | Methane | 0.000 16 |
| Oxygen | 20.94 | Krypton | 0.000 12 |
| Argon | 0.934 | Hydrogen | 0.000 05 |
| Carbon Dioxide | 0.037 | Carbon monoxide | 0.000 01 |
| Neon | 0.0018 | Ozone | 0.000 002 |
| Helium | 0.0005 | Xenon | 0.000 0087 |

**Pollutants in the lower atmosphere**

* Carbon dioxide

**Source:** Burning of fossil fuels

**Problem:** Absorbs radiation which retains heat on Earth

* Carbon monoxide

**Source:** Produced by incomplete combustion in motor vehicles, bushfires

**Problem:** Toxic – absorbed by blood 200x more efficiently than oxygen causing headaches, fatigue, eventual death due to low oxygen transportation.

**Advantage:** Does not build up in the atmosphere due to removal by action of soil organisms or oxidation to carbon dioxide.

* Nitrogen oxides

**Source:** Motor vehicles

**Problem:** Irritant in respiratory system, can be oxidised to form nitric acid and thus acid rain.

* Ozone

**Source:** Reaction of light with nitrogen dioxide forms oxygen free radical

**Problem:** Poisonous in the lower atmosphere, contributes to photochemical smog, oxidises organic tissue disrupting normal biochemical reactions in the body, irritates eyes, causes breathing difficulties, toxic to plants and crops.

**Advantage:** In the stratosphere it absorbs UV-B and UV-B radiation that is harmful to living organisms. Allows UV-A radiation to reach the surface of Earth promoting the production of vitamin-D in the skin.

* Sulfur dioxide

**Source:** Combustion of fuels that contain sulfur (coal and oil) at power stations and metal smelters, volcanoes, bacterial action

**Problem:** Poisonous gas, forms acid rain



**Coordinate Covalent Bonds**

A coordinate covalent bond forms when one atom provides both electrons for the shared pair. Once formed, it is identical to a normal covalent bond.

**Examples:**

* Ammonium ions

Form when ammonia molecules react with hydrogen ions. The non-bonding pair of electrons on the nitrogen atom are shared with the hydrogen ion to form the bond.

* Hydronium ions

The oxygen atom in water has two non-bonding electron pairs. One of these pairs can be shared with a hydrogen ion, forming the hydronium ion.

* Carbon monoxide

This molecule contains a triple bond consisting of two covalent bonds (where the carbon and oxygen both share two electrons each) and one coordinate covalent bond (where the oxygen atom shares two electrons with the carbon atom).

* Ozone

Ozone is a bent molecule like water. It involves two covalent and one coordinate covalent bond.

**Allotropes of oxygen**

**Diatomic Oxygen O2**

* Colourless gas
* Similar density to air
* Less soluble than ozone
* Contains two oxygen atoms linked by a double covalent bond
* Double bond is very strong with a high bond energy 498kj/mol
* Less reactive than ozone
* Supports combustion
* Used in the steel industry to convert iron to steel
* Used in medicine in compressed oxygen bottles to provide patients with oxygen during operations
* Used in space industry as liquid oxygen is used as an oxidiser in the rockets

**Ozone O3**

* Pale blue gas
* 1.5x density of air
* More soluble in water
* Contains three oxygen atoms
* Poisonous gas
* Powerful oxidising agent
* Lower bond energy 445kj/mol thus more reactive than diatomic oxygen
* Forms a free radical oxygen which is extremely reactive
* Used as a bleaching agent in the production of paper
* Used as a disinfectant to kill micro-organisms in water



**Oxygen**

* Oxygen contains 8 electrons
* Configuration 2,6
* Valence shell contains 3 pairs of electrons
* A pair of valence electrons from each atom are shared when O2 is formed
* O2 molecules are split by the absorption of UV light forming O free radicals

**Oxygen Free Radical**

* Contain two pairs of electrons and two unpaired electrons in their valence shell
* The unpaired electrons make the radical very unstable and very reactive
* Exist briefly in lower atmosphere before reacting with other radicals or molecules
* More reactive than ozone
* O free radicals are formed in the thermosphere when UV light causes the decomposition of oxygen molecules

**Formation and Decomposition of Ozone**

Ozone in the stratosphere is constantly being formed and decomposed, converting solar energy to heat energy. Oxygen absorbs UV-C light and photodissociation occurs producing oxygen free radicals which combine with oxygen molecules to produce an energised ozone molecule.

O2 (g) + UV radiation ⇨ 2O (g)

O2 (g) + O (g) ⇨ O3 (g)

The excess energy from the ozone molecule is transferred to another molecule such as nitrogen or oxygen, preventing energised ozone from decomposing.

**O3 (g)**+ N2 (g) ⇨ O3 (g) + **N2 (g)**

Energised ozone molecules decompose back to oxygen unless collisions occur to transfer the excess energy. Collisions are more frequent at lower altitudes, unless the altitude is too low in which case UV light cannot penetrate to initiate the decomposition of oxygen.

Decomposition of ozone can occur in the following ways:

* Ozone absorbs UV-B and UV-C light

The ozone decomposes back into the oxygen free radical and oxygen molecule. Ozone is the only molecule which absorbs UV radiation in this range. The reverse reaction occurs:

O3 (g) + UV radiation ⇨ O2 (g) + O (g)

The oxygen free radicals produced combine with ozone molecules to form oxygen molecules:

O3 (g) + O (g) ⇨ 2O2 (g)

* Ozone decomposes due to other free radicals

Free radicals such as hydroxide, chloride and nitrogen monoxide are found in the upper atmosphere and react with ozone to form the oxygen molecule:

O 3 (g) + NO (g) ⇨ NO2 (g) + O2 (g)

NO2 (g) + O (g) ⇨ NO (g) + O2 (g)

**CFCs and Halons**

Haloalkanes – when an alkane reacts with a halogen (Group 7 element)

**Naming Haloalkanes**

The halogen is used as the prefix, and the number tells us what carbon it is attached to, for example, “1-bromobutane” suggests the bromine is on the first carbon of a butane chain.

The functional groups (halogens) are names in alphabetical order if there is more than one present, for example, “1-bromo, 1-fluoropropane” suggests both the bromine and fluorine is on the first carbon. In the case of having two of the same halogens, we use the di, tri, tertra prefixes as well as denoting the carbon atom each is attached to, for example, “1,2 dichlorobutane” suggests there are two chlorine atoms, one on the first carbon and another on the second carbon.

**Isomers**

Haloalkanes can exist as isomers. The halogens can have variable locations on the carbon chain resulting in a number of different isomers.

***Example 1: C2H2Cl2F2***

***Example 2: C3H5BrCl2***

**Chlorofluorocarbons & Halons**

CFCs: Haloalkane where all hydrogen atoms have been replaced by chlorine and/or fluorine atoms.

Halons: Haloalkane where all hydrogen atoms have been replaced by bromine, chlorine and/or fluorine atoms.

**Origins**

CFCs were originally developed to replace ammonia as a refrigerant gas in refrigerators and air conditioners, and propellants in spray cans as they were unreactive and non-toxic. They were also used to make expanded plastics like polystyrene and as solvents in dry cleaning.

Halons were developed to extinguish fires, specifically in large computer systems and aeroplanes.

**Problems associated with CFCs & Halons**

* Responsible for thinning the ozone layer
* Natural chlorinated compounds (HCl, NaCl) rarely reach the stratosphere (are usually oxidised in the troposphere or dissolve due to high solubility)
* Synthetic halogenated compounds diffuse slowly into the stratosphere and have a long lifetime
* Once present in the stratosphere they decompose on contact with UV light (photodissociation) and to produce reactive chlorine, bromine and fluorine radicals that destroy ozone molecules
* despite most CFCs and halons being inert in the troposphere, once transported to the stratosphere and exposed to UV light, photodissociation occurs creating free radicals
* Restrictions were placed on the use of CFCs as propellants and refrigerants in the early 80s
* The thinning of the ozone layer was discovered in the mid 80s and halogenated hydrocarbons were phased out as alternatives were found (Montreal Protocol 1987)
* 3% of ozone has been lost worldwide in the last 20 years
* Reduction in ozone results in higher UV radiation reaching Earth’s surface
* Increased exposure results in sunburn and skin cancer, damage to plants and sensitive ecosystems such as the Antarctic
* 1% decrease in ozone in the stratosphere equates to a 2% increase in UV radiation reaching Earth’s surface (this can increase skin cancer rates by 4-6%)
* UV radiation can decompose both natural and synthetic polymers
  + Breaks chemical bonds in proteins and DNA
  + Degrades plastics exposed to light for long periods
* UV radiation causes eye cataracts, damage to aquatic organisms, decreases phytoplankton (oxygen producing organisms)and crop productivity

Measures taken to lower the impact of CFCs

* Montreal Protocol/local and international laws/policies
* Banning CFCs
* Alternatives/substitutes

Changes in Ozone Concentrations over Time