Grade 11 Chemistry

Gases and Atmospheric Pressure

Class 16

Mole Fraction

- A dimensionless quantity that expresses the ratio of the number of moles of one component to the number of moles of all components present
- Useful to find the partial pressure of one gas in a mixture of gases

$$X_A = \frac{n_A}{n_{Total}} \qquad P_A = X_A P_{Total}$$

Checkpoint 

A mixture of gases contains 4.46 moles of Ne, 0.74 moles of Ar, and 2.15 moles of Xe. Calculate the partial pressure of the gases if the total pressure is 2.00 atm at a certain temperature.

Pressure of Water Vapour

- You can collect many gases by bubbling them up through water into a container however water vapour gets mixed with the gas sample
- To avoid the error, simply subtract the pressure of the water vapour from the total pressure

$$P_{total} = P_{gas} + P_{H_2O}$$

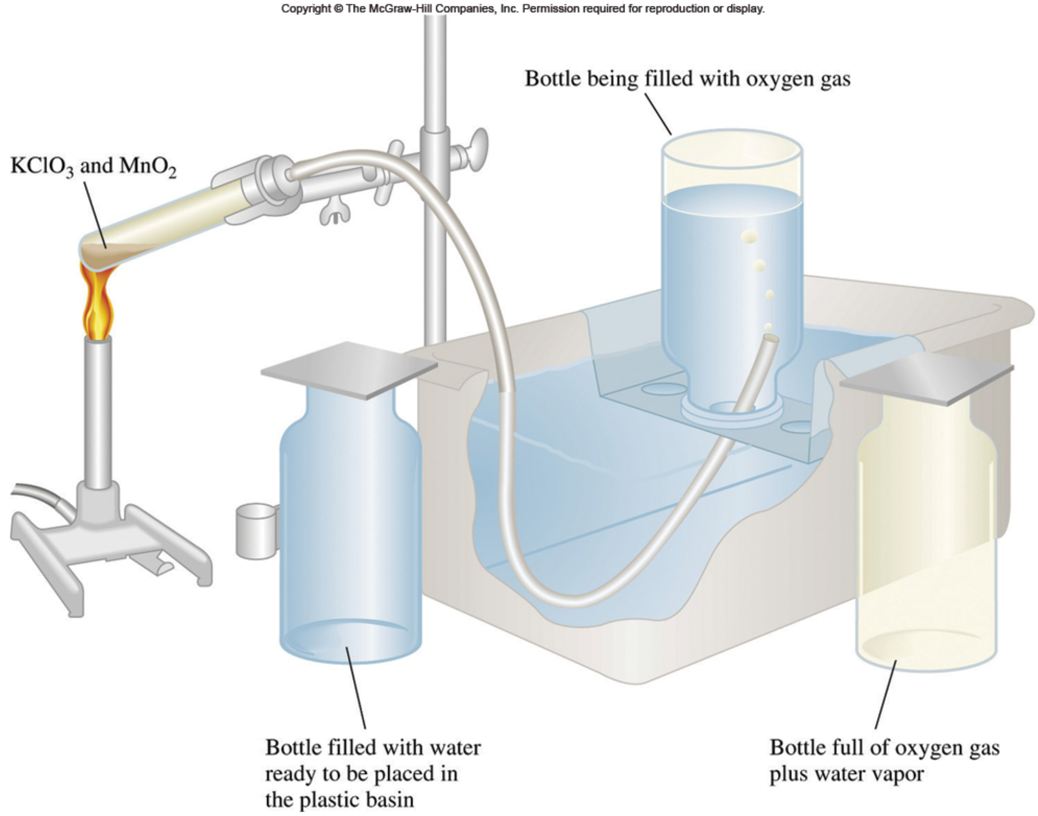


Table 12.3
Pressure of Water Vapour

Temperature (°C)	Pressure (kPa)
17	1.94
18	2.06
19	2.20
20	2.34
21	2.49
22	2.64
23	2.81
24	2.98
25	3.17
26	3.36
27	3.56
28	3.78
29	4.00
30	4.24

- The pressure of water vapour is different for different temperatures
- “Wet Gas” – mixed with water vapour

$$P_{total} = P_{O_2} + P_{H_2O}$$

Checkpoint

A student reacts magnesium with excess dilute hydrochloric acid to produce hydrogen gas. She uses 0.15 g of magnesium metal. What volume of dry hydrogen does she collect over water at 28°C and 101.8 kPa?

Real Gases

- At **high pressure** and/or **low temperatures**, gases no longer behave ideally
- In reality:
  - Particles of a real gas have a significant volume of their own
  - Molecules do attract/repel each other
  - Molecules do not necessarily move in straight lines
  - Collisions are not completely elastic

- At high pressures and/or low temperatures, gases have smaller volumes so the gas molecules are closer together
  - Feel attractive forces
  - The amount of space taken up by the gas molecules is more important if the volume of the container is smaller
- For G11 you will use the ideal gas law but to account for real gases, you will use Van der Waals equation

Van der Waals' Equation

- For real gases:

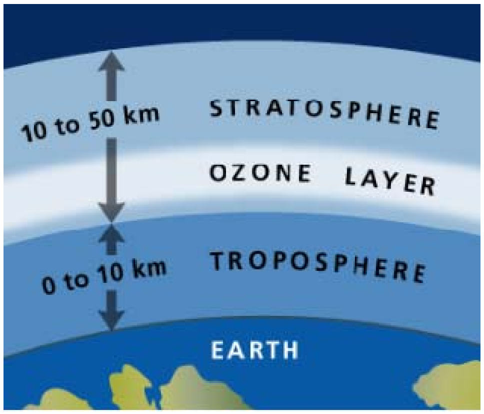
$$\underbrace{\left(P + \frac{an^2}{V^2}\right)}_{\text{corrected pressure}} \underbrace{(V - nb)}_{\text{corrected volume}} = nRT$$

- Where "a" and "b" are constants

TABLE 5.4
van der Waals Constants of Some Common Gases

Gas	a $\left(\frac{atm \cdot L^2}{mol^2}\right)$	b $\left(\frac{L}{mol}\right)$
He	0.034	0.0237
Ne	0.211	0.0171
Ar	1.34	0.0322
Kr	2.32	0.0398
Xe	4.19	0.0266
H_2	0.244	0.0266
N_2	1.39	0.0391
O_2	1.36	0.0318
Cl_2	6.49	0.0562
CO_2	3.59	0.0427
CH_4	2.25	0.0428
CCl_4	20.4	0.138
NH_3	4.17	0.0371
H_2O	5.46	0.0305

The Ozone Cycle



$$O_3 \overset{UV}{\leftrightarrow} O_2 + O$$

- Ozone, O_3 absorbs UV radiation from the Sun which separates the ozone into oxygen gas and an oxygen atom
- After a few more steps, ozone is reformed
- This cycle allows UV radiation to be absorbed preventing them from causing cancer
- Ozone naturally exists in the stratosphere (25km above Earth)

Pollutants and Ozone

- Gases from living and non-living processes on Earth's surface interact with the gases in the atmosphere
- Chlorofluorocarbons (CFCs) are an important class of polluting gases; they are stable near the ground but when they get into the atmosphere, they interfere with the ozone cycle
- CFCs are found in refrigerants, foam products and aerosol devices

Chlorofluorocarbons

- By 1974, 1.5 million tonnes of CFCs had been produced in the U.S.
- Because they are so stable, what happens when they are exposed to high levels of radiation in the atmosphere?
- In 1985, scientists noticed a hole in the ozone layer above the Antarctic

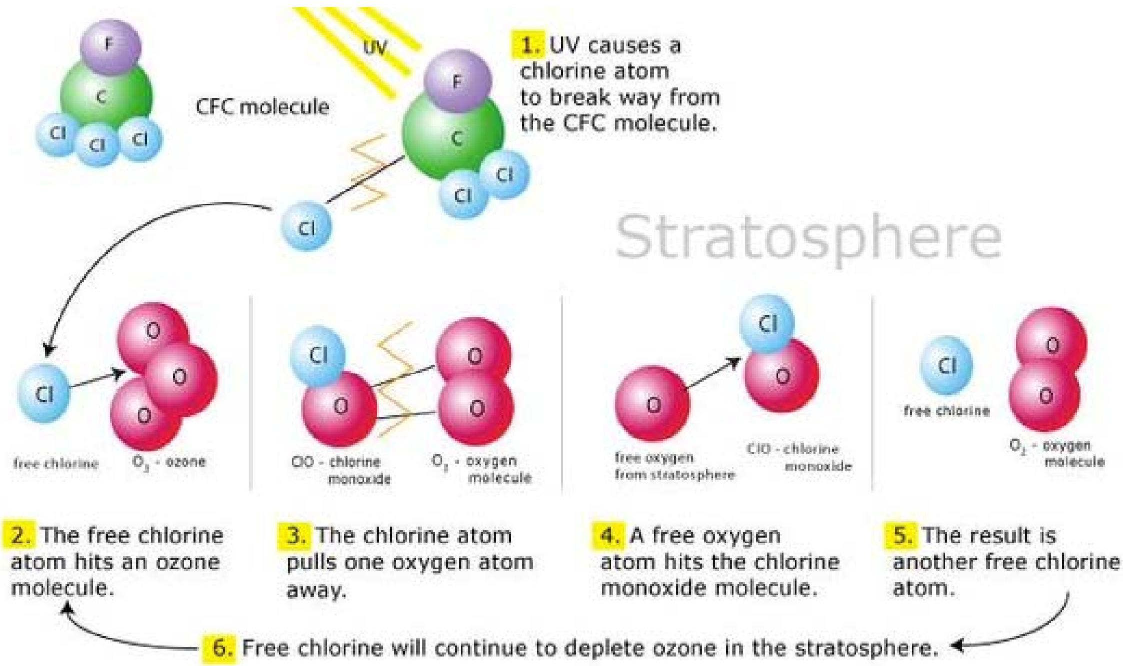
- CFCs in the atmosphere break apart under radiation to produce chlorine atoms which destroy ozone molecules

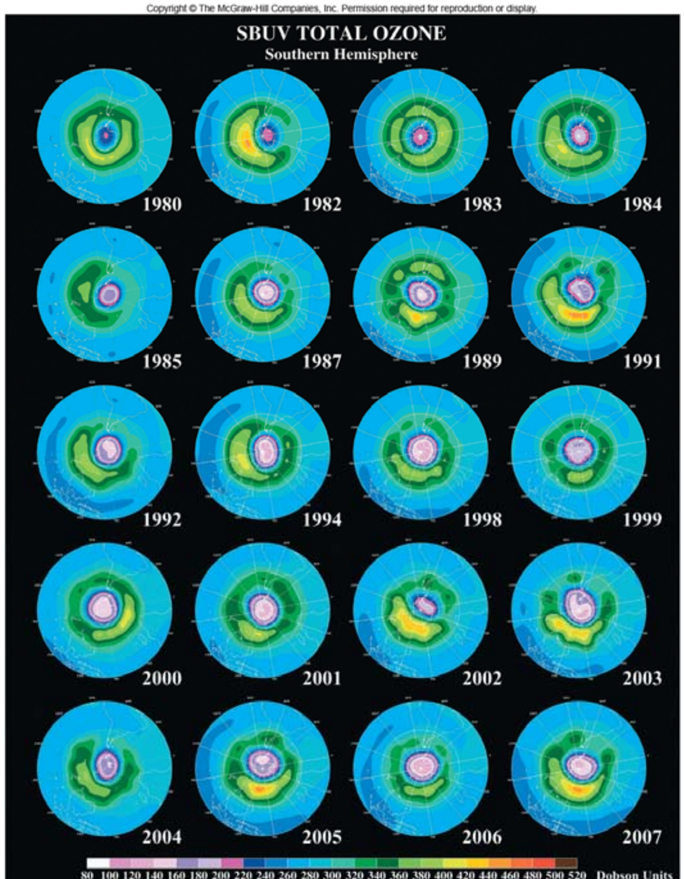
$$Cl + O_3 \rightarrow ClO + O_2$$

- The ClO reacts with an oxygen atoms and releases the chlorine atom which can continue the cycle

$$ClO + O \rightarrow Cl + O_2$$

- 1 chlorine atom can destroy thousands of ozone molecules





NOAA

The depletion of the ozone is shown in purple over the South Pole

Effects of Ozone Depletion

- Reduced ozone levels means that more UV radiation from the Sun reaches the Earth
- Can cause:
  – Skin Cancer
  – Eye damage
  – Phytoplankton damage

Smog

- Ozone is helpful in the stratosphere but when it is in the troposphere, it is considered smog
- Ozone is a pollutant with a harsh odour and can cause respiratory problems, prevents plant growth, reduces the productivity of crops and damages forests

Improving Air Quality

- We now use environmentally friendly refrigerants and air-conditioners
- Used CFCs are being recycled
- Montreal Protocol in 1987 reduced the use of CFCs worldwide – all production and consumption of substances that deplete the ozone layer would be phased out by the year 2000 in developed countries
- Includes: CFCs, CCl_4, methyl chloroform, methyl bromide
- We hope that the ozone layer will recover within 50-60 years