SUPPLEMENTARY TOPIC 2
MORE ABOUT GASES.

Behaviour of gases.
From the kinetic molecular theory (Supplementary Topic 1) the molecules of a gas are, on average, relatively far apart and in constant motion in straight lines between collisions, displaying considerable rapid translational motion. When the molecules collide, they rebound with no loss of energy. Also according to this model, collisions of these fast-moving molecules with the walls of the container are the cause of pressure. In this Topic, properties of gases are further examined with regard to their response to pressure applied and to temperature, their solubility in water and their molar volumes.

Atmospheric pressure.
At sea level, the atmosphere of earth causes a pressure of 1.0 kg to be exerted on each square centimetre of any object (written as 1.0 kg cm.$^{-2}$). We are oblivious to this pressure, just as a fish is not aware of the much larger pressures to which it is subjected in the sea. Without that constant pressure, the human body cannot function properly. For example, the action of the lungs in breathing depends on air pressure causing fresh air to flow into the lungs when muscles expand the rib cage. In space, a person not maintained at atmospheric pressure by a space suit would experience the boiling of all water in his body at body temperature due to the vacuum of space. Rapid travel up or down a steep hill may cause pressure imbalance on either side of the ear drum, leading to temporary loss of hearing.

Effect of pressure on the volume of a gas.
When pressure is increased on a gas trapped in a suitable container, its volume decreases. For example, if one blocks off the outlet of a bicycle pump and then tries to push the piston down the pump, the air in the barrel of the pump is compressed. Releasing the handle allows the gas to expand back to its original volume (assuming none leaked out!). Further, as the piston is pushed down the barrel, increasing resistance is felt, until ultimately more force than one can supply is needed to compress the air. Careful observations of the relationship between the pressure applied to a gas and its volume have been made using gases trapped in a glass column. The pressure is conveniently supplied by a mercury reservoir which can be placed at measured heights.

These experiments show that provided the temperature remains unchanged, the volume of a trapped gas is inversely proportional to the applied pressure and this relationship is known as BOYLE'S LAW. This means that, as pressure is increased
on the gas, its volume decreases proportionately. For example, if the pressure (P) were doubled, the volume (V) would be halved. Boyle's law can be expressed by the equation

\[ P \times V = \text{constant} \quad \text{or} \quad P_1 \times V_1 = P_2 \times V_2 \]

where \( V_1 \) and \( V_2 \) are the volumes when the pressures are \( P_1 \) and \( P_2 \) respectively. As an example, 1.0 L of air at 1.0 atmosphere pressure at a constant temperature would occupy 0.5 L at 2.0 atmospheres as \( 1.0 \times 1.0 = 0.5 \times 2 \).

The accompanying diagrams illustrate how the volume of a gas changes as the pressure increases.

In terms of the kinetic molecular theory presented in Supplementary Topic 1, the reduced volume available to the gas results in more collisions between the gas molecules and the container walls, leading to increased pressure, as illustrated in this diagram.

Gases under pressure have the ability to do work when they are allowed to expand again, due to reduction of the pressure, thus releasing the stored energy in the compressed gas. Examples include the use of air compressors to drive jack hammers, and the firing of pellets in an air rifle.

**Check your understanding of this section.**

Calculate the pressure required to compress air which occupies a volume of 1.00 L at 1.00 atmospheres to 0.10 L at the same temperature.
Effect of temperature on the volume of a gas.
When a gas in a fixed volume container is heated, the pressure inside the container increases as the gas cannot expand. If the container is one which maintains constant pressure instead of constant volume, (e.g. a piston in a cylinder), the volume (V) expands as temperature (T) increases. A graph of the volume plotted over a range of temperatures is found to be a straight line, as illustrated on the next page. If this line were to be extrapolated to zero volume, the temperature would be –273°C. As no sample of matter could reduce to zero volume, this temperature represents the lowest theoretical temperature possible.

A temperature scale called the **KELVIN SCALE** adopts this temperature as its zero, sometimes called "**ABSOLUTE ZERO**", and uses the same interval for a degree as is used for the Celsius or centigrade scale. Thus zero on the Celsius scale would be 273 on the Kelvin scale, shown as 273 K (no degrees sign is used). Likewise, the typical room temperature in summer of 25°C is 298 K. If the Kelvin scale were used, the volume of a gas held at constant pressure is **DIRECTLY PROPORTIONAL** to the temperature. i.e. \[ V = \text{Constant} \times T, \quad \text{or} \quad \frac{V_1}{T_1} = \frac{V_2}{T_2} \] for two temperatures \( T_1 \) and \( T_2 \). This relationship is referred to as **CHARLES'S LAW**. The following diagrams illustrate how the volume and pressure of a contained gas respond to temperature changes.
As an example, 1.0 L of air at 25°C (298 K) expands when heated to 100°C (373 K) so that the final volume is $1.0 \times \frac{373}{298} \text{ L} = 1.25 \text{ L}$ at the same pressure.

The volume, pressure and temperature relationships for gases given by the laws of Boyle and Charles can be combined into a single equation. For two sets of conditions 1 and 2,

$$\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$$

where $T$ is in Kelvin.

**Check your understanding of this section.**

Calculate what temperature 1.0 L of air at 273 K would have to be heated to in order that its volume increased to 2.0 L at the same pressure.

**Kinetic molecular explanation of Charles’s law.**

The kinetic molecular model of gases explains Charles's Law in terms of the molecules moving faster at higher temperatures and thus having more energy. Consequently collisions between the molecules and the walls of the container occur more frequently and result in more force on the walls - i.e. an increase in pressure occurs. If the container walls are not rigid, then according to Boyle's Law, the volume will expand. The effect of temperature on gas pressure and volume is seen in the increase in pressure observed when the tyres of a motor car become hotter as the car is driven. The tyres are able to increase their volume only slightly, leading to increased pressure at the higher temperature as in the above example. Hence one needs to check the pressures when the tyres are cold to avoid under-inflation. The kinetic molecular theory’s explanation of Charles’s Law showing the relation between volume and temperature at constant pressure is illustrated in the following diagrams.
Units used for pressure.
There are numerous units used to measure the pressure of gases. The most commonly
used are atmospheres, millimetres of mercury and Pascals. As atmospheric pressure
varies constantly with weather conditions and also with altitude, it is necessary to define
one atmosphere of pressure as 760 millimetres of mercury (mmHg). This refers to the
height of a column of mercury that can be supported in a glass tube in an instrument
called a barometer. The Pascal (Pa) is the SI unit of pressure and one atmosphere
corresponds to 101.3 kilopascals (kPa).

Example 1.
A motor car tyre at 25.0°C (298K) has a volume of 50.0 litres at a pressure = 240
kilopascals. After travelling on a hot road, the tyre volume is 50.5 litres and the
temperature of the air inside is 70.0°C (343K). Calculate the pressure now in the tyre.

\[ P_1 = 240 \text{ kPa}, \quad V_1 = 50.0 \text{ L}, \quad T_1 = 298 \text{ K} \]
\[ P_2 = ?, \quad V_2 = 50.5 \text{ L}, \quad T_2 = 343 \text{ K} \]

Substituting in the combined Boyle’s and Charles’s law equation,

\[
\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}
\]

\[
\frac{240 \times 50.0}{298} = \frac{P_2 \times 50.5}{343}
\]

\[ P_2 = 274 \text{ kPa} \]

The ideal gas equation.
The combined laws of Boyle and Charles given previously show that \( PV/T \) for any fixed
amount of gas is a constant provided that \( T \) is expressed in the Kelvin temperature scale.
If an amount of gas equal to exactly 1 mole is used, this constant is called the
universal gas constant, symbolised as \( R \). The numerical value of \( R \) will
depend on the units used to express the pressure and volume of the gas. If the pressure is
measured in atmospheres and the volume measured in litres (remembering that \( T \) must
always be in Kelvin), then regardless of which gas is taken, the constant \( R \) has the value
0.0821 litre atmospheres per Kelvin mole. It can be shown that pressure \( \times \) volume has
energy units and so if the SI derived units for \( P \) (kilopascals) and volume (litres) are
used, then the resulting energy units are in the SI system, Joules. Using kilopascals and
litres, the value and units for \( R \) become 8.31 joules per Kelvin mole (written as \( J \text{ K}^{-1} \text{ mol}^{-1} \)). For \( n \) moles of gas, the equation becomes

\[
PV = nR \quad \text{or} \quad PV = nRT
\]
This is called the **Ideal Gas Equation** but it applies even to real gases subject to the limits given previously where pressures are not extremely high or temperatures extremely low.

One implication of this equation is that the volume of exactly one mole of any gas at a specified temperature and pressure is always the same. For example, at “room temperature and pressure” which are generally taken as 25°C (298 K) and one atmosphere (101.3 kPa), the **Molar Volume** of any gas is 24.5 L.

**Example 2.**
The reaction between oxygen and sulfur dioxide is shown in the following equation:

\[ \text{O}_2(g) + 2\text{SO}_2(g) \rightarrow 2\text{SO}_3(g) \]

(a) Starting with 1.00 mole of oxygen and 2.00 mole of sulfur dioxide, calculate the volume of sulfur trioxide produced at 25°C and 101.3 kPa assuming complete reaction.

From the equation, one mole of oxygen produces two moles of sulfur trioxide.

Molar volume of any gas at 25°C and 101.3 kPa = 24.5 L.

∴ volume of \( \text{SO}_3(g) \) at these conditions = \( 2 \times 24.5 \text{ L} = 49.0 \text{ L} \)

(b) Calculate the volume of this amount of sulfur trioxide at 50°C and 200.0 kPa.

Either use the combined Boyle’s and Charles’s law equation or use the ideal gas equation.

Either method requires the temperature to be in the Kelvin scale.

\[ 50°C = (273 + 50) \text{ K} = 323 \text{ K} \]

Using \( \frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2} \)

\[ \frac{101.3 \times 49.0}{298} = \frac{200.0 \times V_2}{323} \]

\[ V_2 = 26.9 \text{ L} \]

Using the ideal gas equation, \( PV = nRT \),

\[ 200.0 \times V = 2.00 \times 8.314 \times 323 \]

∴ \( V = 26.9 \text{ L} \)
Mixtures of gases.
If a mixture of gases is in an enclosed container, then each gas expands to uniformly and independently fill the container and the PARTIAL PRESSURE of each gas is the pressure which it would exert if that gas alone occupied the container. The total pressure is the sum of all the partial pressures of all the component gases in the mixture.

For a mixture of gases 1, 2, 3..... \[ P_{\text{total}} = P_1 + P_2 + P_3 + ..... \]

Solubility of gases in liquids.
In earlier Topics, it was shown how ionic solids such as sodium chloride dissolve in water because the ions combine with the polar water molecules, releasing enough energy to break the ionic crystal lattice. Gases also dissolve in liquids such as water, but usually they do not produce ions in solution. For most gases, instead of the relatively strong attractive forces existing between ions and water molecules, there are much weaker attractions which allow at least some of the gas to transfer into the dissolved state. For example, the dissolution of oxygen gas in water could be represented by the equation

\[ \text{O}_2(g) \rightarrow \text{O}_2(aq) \]

where \( \text{O}_2(aq) \) represents the association between water molecules and the oxygen molecule. The amount of gas present in a given solution depends on the following three factors.

(i) the particular gas
(ii) the temperature
(iii) the pressure of gas over the solution.

(i) Some gases interact to a significantly greater extent with the solvent than do other gases. For example, the gas ammonia, \( \text{NH}_3 \), is very soluble in water because it can hydrogen bond to the water molecules (Topic 13) and also it forms ions of \( \text{NH}_4^+ \) and \( \text{OH}^- \) to a small extent. By comparison, gases such as oxygen and nitrogen are much less soluble in water because they have no hydrogen bonding with the solvent and do not form ions in water. Hydrogen chloride is very soluble in water due to the reaction to form hydrochloric acid which arises from the complete ionization of hydrogen chloride to \( \text{H}^+ \) and \( \text{Cl}^- \) ions as shown in the following equation.

\[ \text{HCl}(g) \rightarrow \text{H}^+(aq) + \text{Cl}^-(aq) \]

However, generally gases do not form ions or hydrogen bond with the solvent.

(ii) The solubility of a gas in any solvent decreases as the temperature increases. If water is boiled for long enough, all dissolved gases will be expelled. However, if the boiled water is then left exposed to the air for a while, gases from the atmosphere will again dissolve until their concentrations return to the values that existed before the water was boiled. Once this occurs, a dynamic equilibrium is established between gas dissolving and gas escaping from the solution. The reduction in gas solubility with increased
temperature is due to the increased energy of the dissolved gas molecules which allows them to escape from the solvent.

\[ \text{e.g. } \text{O}_2(\text{aq}) + \text{heat} \rightarrow \text{O}_2(\text{g}) \]

The reduction in solubility of a gas with increased temperature is the opposite effect to that observed for most solids dissolving. The reason is that increasing the temperature imparts increased energy to the dissolved gas molecules which can then escape from the dissolved state to the gaseous state more easily, while for solids, the additional energy imparted at a higher temperature allows its component ions or molecules to escape from the crystal lattice more easily and thus go into the dissolved state.

The reduced solubility of gases in solvents at higher temperatures has a drastic effect on aquatic life. Reduced oxygen content in water due to heating (e.g. by discharge of hot water from power station cooling towers) can cause the death of fish and other oxygen-breathing aquatic animals. Conversely, cooler ocean temperatures encourage marine life, for example the Humboldt current that flows north along the coast of South America.

The following table gives the solubilities of a number of gases at 1 atmosphere pressure and a temperature of 0°C in mole per litre. The last two gases in the list have much stronger interactions with water due to ionisation and/or hydrogen bonding.

<table>
<thead>
<tr>
<th>Gas</th>
<th>Solubility (mol/L)</th>
</tr>
</thead>
<tbody>
<tr>
<td>nitrogen</td>
<td>$1.05 \times 10^{-3}$</td>
</tr>
<tr>
<td>oxygen</td>
<td>$1.3 \times 10^{-4}$</td>
</tr>
<tr>
<td>hydrogen</td>
<td>$9.6 \times 10^{-4}$</td>
</tr>
<tr>
<td>carbon monoxide</td>
<td>$1.57 \times 10^{-3}$</td>
</tr>
<tr>
<td>hydrogen sulfide</td>
<td>0.208</td>
</tr>
<tr>
<td>ammonia</td>
<td>11</td>
</tr>
<tr>
<td>hydrogen chloride</td>
<td>12</td>
</tr>
</tbody>
</table>

(iii) Increased pressure of a gas above a solvent causes more gas to dissolve. This effect is illustrated in the following diagram.
Henry's Law states that the solubility of a gas in a solvent is directly proportional to the pressure of the gas above the solution. Just as solutions of solids in liquids reach saturation when sufficient solid has been added, so too do solutions of gases become saturated when no more of the gas can be made to dissolve, no matter how high the pressure used. The amount of gas that can dissolve before saturation is reached at a given temperature will therefore increase as the pressure of the gas over the solution is increased. If a mixture of gases is present over the liquid, then the amount of each dissolving is proportional to its partial pressure over the solution.

Examples of Henry's Law in action include the "divers bends" and the production of soda water in a soda siphon. Divers bends result from nitrogen which has dissolved to a greater extent in the blood at the higher pressures supplied in order to balance the pressure of water at great depths. While the diver stays deep down in the water, the nitrogen remains dissolved, but if the diver surfaces rapidly the nitrogen solubility falls to match the external pressure. The excess nitrogen comes out of solution in the blood as bubbles which cause considerable pain and possibly death due to blockage of blood vessels. The solubility of oxygen in the blood would also be reduced as the diver surfaced, but it can be removed in the body by respiration. Carbonated drinks are made by forcing carbon dioxide to dissolve at high pressure in water. While the drink remains sealed, the gas stays in solution but when the pressure is released by removing the lid, the carbon dioxide comes out of solution rapidly and forms the familiar bubbles associated with fizzy drinks. The process of dissolution of a gas in water is hastened by increasing the surface area of water in contact with the gas. This is the reason for bubbling air through the water in an aquarium tank in order to increase the amount of oxygen available.

Objectives of this Topic.
When you have completed this Topic, you should have achieved the following goals:

1. Understand that all objects are subject to atmospheric pressure which depends on height above sea level.

2. Know the qualitative and quantitative relationship between the volume of an enclosed gas and its pressure at a fixed temperature.

3. Know the qualitative and quantitative relationship between the volume of a gas and its temperature on the Kelvin scale.

4. Understand how the Kelvin temperature scale is derived and be able to convert Celsius and Kelvin temperatures.

5. Be able to apply the quantitative relationships from Boyle's law and Charles's law to some simple calculations.
6. Know that the amount of gas that dissolves in a liquid is determined by the temperature of the liquid, the pressure of gas over the liquid and characteristics of each individual gas.

7. Know that in a mixture of gases in an enclosed container, each gas expands to uniformly fill the container and exerts its partial pressure which is the pressure it would exert if that gas solely occupied the container.

8. Know that the total pressure in a mixture of gases is the sum of the partial pressures of all the component gases.

9. Know that gases dissolve in liquids as a result of interactions between the gas molecules and the solvent molecules.

10. Be able to differentiate between the effects of increased temperature on the solubility of solids in liquids and gases in liquids.

11. Be able to explain the gas laws in terms of the kinetic molecular model.

12. Be able to use the ideal gas equation in computations involving volumes, pressures, temperatures and amounts of gases.

**Recommended follow up chemical modules:**

Section: General Chemistry
Module: Behaviour of Gases
Topic: The gas laws; ideal gas equation; real gases.

**SUMMARY.**

Gases consist of particles (atoms or molecules) which are in rapid motion, colliding with each other and with the boundaries of their containers. One consequence of these collisions is the existence of air pressure. At sea level, the pressure due to the atmosphere is largely unnoticed but very significant. The volume, temperature and pressure of a contained gas are related by two laws: Boyle's law which states that the volume of gas is inversely proportional to the applied pressure at a fixed temperature and Charles's law which states the volume is directly proportional to the absolute temperature at a fixed pressure. These laws can be embodied in a single relationship, viz $P \propto V/T$ (T in Kelvin) for any enclosed gas is a constant.
The relation of volume to temperature of a gas leads to the proposal of the absolute or Kelvin temperature scale which is based on a zero which would correspond to the temperature at which the gas would theoretically have zero volume. This temperature is $-273^\circ C$, and a one degree interval is the same as on the Celsius scale.

The ideal gas equation relates the laws of Boyle and Charles to the amount of gas present under the prevailing conditions by including the constant R. That equation is $PV = nRT$ and if pressure is specified as kilopascals, volume as litres and temperature as Kelvin, then $R$ has the value 8.314 joules per Kelvin mole. The ideal gas equation is subject to the same limitations as the laws of Boyle and Charles but is a satisfactory representation of real gas behaviour provided the temperature is not very low or the pressure extremely high.

In a mixture of contained gases, each gas expands uniformly and independently to occupy the total volume of the container and each gas exerts its own partial pressure. The sum of the partial pressures of all the gases in the mixture is equal to the total gas pressure.

The solubility of gases in liquids such as water decreases at higher temperatures (due to the molecules having more energy to escape from the solution), increases with increased pressure of the gas over the solution, and also depends on factors such as the structure of the gas molecules and how they interact with the solvent.

**TUTORIAL QUESTIONS - SUPPLEMENTARY TOPIC 2.**

1. Why is it necessary to pressurise the cabins of high flying aircraft?

2. What does the barometer measure, and why does its reading vary from day to day?

3. How is the Kelvin temperature scale defined? What would $20^\circ C$ be on the Kelvin scale?

4. Air (1.0 L) at 1.0 atmosphere pressure and $25^\circ C$ has the applied pressure increased to 10 atmospheres at the same temperature. What would be the new volume observed?

5. What would be the volume of the same sample of gas in the previous question if the temperature were to be increased to $300^\circ C$, the pressure remaining at 1.0 atmosphere?

6. Why may the discharge of water used for cooling from industrial plants pose problems for waterways?

7. A glass of water taken from the tap in winter often forms many small bubbles throughout the liquid upon standing in a warm room. Explain this observation.
8. Explain how the expansion of gases with increased temperature is relevant to the operation of the internal combustion engine.


10. What causes the "bends" suffered by divers who surface quickly from significant depths of the ocean? How does treatment in compression chambers cure them?

11. Why is the gas hydrogen chloride much more soluble in water than chlorine or hydrogen gases?

12. The solubility of an ionic compound such as copper(II) sulfate increases if the temperature of the solution is raised but the solubility of a gas such as carbon dioxide is reduced at higher temperatures. Explain why this is so.


14. What will be the volume of a sample of carbon dioxide at 98.0 kPa if its volume at 96.4 kPa is 223 mL, the temperature being unchanged?

15. What volume will be occupied by a sample of hydrogen at 153 kPa if the volume at 98.6 kPa is 325 mL, the temperature being held constant?

16. A sample of hydrogen sulfide was collected in a 250 mL flask at a pressure of 98.6 kPa and 310 K. What volume would the gas occupy at 369 kPa and 313 K?

17. A mixture of Ar (3.28 x 10⁻³ mole) and N₂ (1.92 x 10⁻² mole) was collected in a vessel whose capacity was 62.5 mL. What would be the total gas pressure at 298 K?

ANSWERS TO TUTORIAL SUPPLEMENTARY TOPIC 2

1. The atmospheric pressure outside of high flying aircraft is considerably less than the one atmosphere that we normally experience, so unless the interior pressure is maintained at close to one atmosphere, there would be insufficient oxygen to support life. Sudden depressurisation in aircraft is not unknown and in this situation, oxygen masks drop from the ceiling.

2. A barometer measures atmospheric pressure. Air pressure varies due to areas of high and low pressure regions in the atmosphere that constantly circle the earth as seen on the weather maps. This constant variation results from heating effects from the sun and oceans and the motion of the earth.
3. The Kelvin temperature scale is defined by zero Kelvin being the temperature at which, according to Charles’s Law, the volume of an ideal gas would be zero and this corresponds to –273 degrees Celsius. [This is unattainable as the gas could not disappear.] A one degree interval on the Kelvin scale is defined as being the same as for the Celsius scale. To convert a temperature expressed on the Celsius scale to Kelvin, add 273. Thus 20°C would be (273 + 20) K = 293 K. Note that no degree sign is used when temperatures are expressed as Kelvin.

4. From Boyle’s law, \( P_1 \times V_1 = P_2 \times V_2 \), (noting that \( T \) is constant), substitute the two sets of conditions 1 and 2.

\[
1.0 \times 1.0 = 10 \times V_2 \\
\therefore V_2 = 0.10 \text{ L}
\]

5. From Charles’s law, \( \frac{V_1}{T_1} = \frac{V_2}{T_2} \), (noting that \( P \) is constant), substitute the two sets of conditions 1 and 2 where \( T \) is in Kelvin.

\[
1.0 / 298 = V_2 / (300 + 273) = V_2 / 573 \\
\therefore V_2 = 1.92 \text{ L}
\]

6. The solubility of gases decreases as the water temperature increases, so less oxygen would be available for marine organisms. In extreme cases, aquatic life dies due to there being insufficient dissolved oxygen as a result of the high water temperature.

7. Water from the tap contains gases dissolved from the atmosphere, mostly oxygen and nitrogen. As the water temperature rises to room temperature, some of the gases may exceed their maximum solubility for the conditions of their partial pressure and temperature, causing them to come out of solution rapidly as bubbles.

8. A mixture of air and petrol vapour is ignited in the cylinder of the internal combustion engine when the piston is at the top of the cylinder. The heat released by the combustion causes the mixture of gases in the cylinder to expand and increase the pressure in the cylinder. This leads to the piston being forced down the cylinder. The motion of the pistons is transferred to a crankshaft which ultimately leads to the motion of the car’s wheels.

9. According to the kinetic molecular theory, gases contain rapidly moving particles (molecules or atoms) which are colliding with each other and the walls of any container in which the gas is trapped. The pressure observed for a trapped gas is
the result of the particles colliding with the walls of the container. The more collisions that are occurring, the greater will be the observed pressure, so if the volume is reduced as in a piston/cylinder system, then there will be more collisions between the particles and the walls of the container and thus an increase in pressure. If the temperature is kept constant so that the particles do not change their energy, this leads to Boyle’s law, P × V = a constant. The theory also postulates that the energy of the particles is related to the temperature, so increasing the temperature causes there to be more collisions with higher energy. This translates to an increase in volume if the gas is in a constant pressure environment or to an increase in pressure if the gas is trapped in a constant volume system. If the temperature is measured in Kelvin, the relation V / T = a constant results which is Charles’s law.

10. Divers are subjected to the external pressure of the water at the diving depth and in order to breathe, air must be supplied at a pressure to balance the water pressure. From Henry’s law, the higher air pressure leads to more oxygen and nitrogen molecules dissolving in the bloodstream. If the diver surfaces quickly, the dissolved gases, mainly the nitrogen, come out of solution and form bubbles which can damage nerves and blood vessels. The result is extreme pain and even death. To avoid the bends, divers surface at a controlled rate with stops for periods at gradually decreasing depths in order to give enough time for the gases to be expelled via the lungs.

11. Hydrogen chloride gas consists of covalent molecules but it reacts vigorously with water to form H⁺(aq) and Cl⁻(aq) ions, a solution called hydrochloric acid which is a strong acid because it is totally ionised. This process allows much more HCl gas to dissolve compared to the amount that would dissolve without the reaction taking place. Hydrogen and chlorine gases do not react with water to form ions in solution and so without this mechanism, their solubilities in water are considerable less than that of hydrogen chloride.

12. Ionic solids dissolving in water require energy for the ions to break away from the attraction of the crystal lattice, so higher temperatures allow more ions to escape and thereby increase the solubility. Gases dissolved in liquids experience an attraction to the solvent molecules which maintains them in the dissolved state. However, as more energy is supplied by heating, the dissolved gas molecules overcome the attraction to the solvent and escape more readily from the solution. Boiling a solution of a gas such as carbon dioxide will ultimately expel all of the CO₂ molecules from the solution.

13. Henry’s law states that at a given temperature the solubility of a gas in a liquid is proportional to the pressure of that gas above the liquid. Examples of this law in
action include carbonated drinks where carbon dioxide is forced under pressure to
dissolve in water and the dissolution of oxygen in water which is essential for all
aerobic aquatic life.

14. \[ PV = \text{constant or } P_1 \times V_1 = P_2 \times V_2 \]
   
   \[ P_1 = 96.4 \text{ kPa and } V_1 = 223 \text{ mL} \]
   \[ P_2 = 98.0 \text{ kPa and } V_2 = ? \]
   
   So \[ 96.4 \times 223 = 98.0 \times V_2 \]
   
   \[ V_2 = (96.4 \times 223) / 98.0 = 219 \text{ mL} \]
   
   [Note: the units of \( P \) and \( V \) must be the same on both sides of the equation.]

15. As temperature is constant, Boyle's Law can be used.
   \[ PV = \text{constant or } P_1 \cdot V_1 = P_2 \cdot V_2 \]
   
   \[ P_1 = 98.6 \text{ kPa and } V_1 = 325 \text{ mL} \]
   \[ P_2 = 153 \text{ kPa and } V_2 = ? \]
   
   So \[ 98.6 \times 325 = 153 \times V_2 \]
   
   \[ V_2 = (98.6 \times 325) / 153 = 209 \text{ mL} \]

16. The temperature and pressure both change so the combined Boyle's and Charles's
    Law expression must be used. \[ P_1 \cdot V_1 / T_1 = P_2 \cdot V_2 / T_2 \]
   
   \[ P_1 = 98.6 \text{ kPa, } V_1 = 250 \text{ mL} \text{ and } T_1 = 310 \text{ K (Note that temperature must be} \]
   
   \[ \text{expressed in Kelvin)} \]
   
   \[ P_2 = 369 \text{ kPa, } V_2 = ? \text{ and } T_2 = 313 \text{ K} \]
   
   So \[ (98.6 \times 250) / 310 = (369 \times V_2) / 313 \]
   
   \[ V_2 = (98.6 \times 250 \times 313) / (310 \times 369) = 67.4 \text{ mL} \]

17. The Ideal Gas Equation is required to solve this problem.
   \[ PV = nRT \]
   
   The constant \( R \) has a value which depends on the units for \( P \) and \( V \). (Units for \( n \)
   
   are always moles).
   
   If \( P \) is expressed in kPa and \( V \) in litres, \[ R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}. \]
   
   Temperature is in \( \text{Kelvin) } \]
   
   \[ P = ? \text{ kPa, } V \text{ 62.5 mL} = 0.0625 \text{ L} \]
   
   and \( n \) = total moles of gas present = \[ 3.28 \times 10^{-3} + 1.92 \times 10^{-2} \]
   
   \[ = 2.248 \times 10^{-2} \text{ mol} \]
   
   \[ R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1} \text{ and } T = 298 \text{ K (Note: } T \text{ must be in Kelvin)} \]
   
   \[ P \times 0.0625 = 2.248 \times 10^{-2} \times 8.314 \times 298 \]
   
   \[ P = (2.248 \times 10^{-2} \times 8.314 \times 298) / 0.0625 \]
   
   \[ = 891 \text{ kPa} \]