Marks 5

- Calcium carbide, CaC_2, reacts with water to produce a gas and a solution containing OH^- ions. A sample of CaC_2 was treated with excess water and the resulting gas was collected in an evacuated 5.00 L glass bulb. At the completion of the reaction, the pressure inside the bulb was 1.00×10^5 Pa at a temperature of 26.8 °C. Calculate the amount (in mol) of the gas produced.

5.00 L corresponds to 5.00×10^{-3} m^3 and 26.8 ^{o}C corresponds to (26.8 + 273.0) K = 299.8 K. Using the ideal gas law:

$$PV = nRT$$

$$n = PV / RT$$
$$= (1.00 \times 10^5 \text{ Pa})(5.00 \times 10^{-3} \text{ m}^3) / ((8.314 \text{ Pa m}^3 \text{ mol}^{-1} \text{ K}^{-1})(299.8 \text{ K}))$$
$$= 0.201 \text{ mol}$$

Answer: **0.201 mol**

Given that the mass of the gas collected was 5.21 g, show that the molar mass of the gas is 25.9 g mol^{-1}.

As the number of moles = mass / molar mass:

molar mass = mass / number of moles = 5.21 g / 0.201 mol = 25.9 g mol^{-1}

Suggest a molecular formula for the gas and write a balanced equation for the reaction that occurred.

As the gas was produced from CaC_2 which contains C_2^{2-}, a likely formula is C_2H_2:

molar mass = (2 × 12.01 (C) + 2 × 1.008 (H)) g mol^{-1} = 26.0 g mol^{-1}

This is formed by addition of H_2O:

$$CaC_2(s) + 2H_2O(l) \rightarrow Ca^{2+}(aq) + 2OH^-(aq) + C_2H_2(g)$$

Marks 3

- A sample of gas is found to exert a pressure of 7.00×10^4 Pa when it is in a 3.00 L flask at 10.00 °C. Calculate the new volume if the pressure becomes 1.01×10^5 Pa and the temperature is unchanged.

Using the ideal gas law, $PV = nRT$, the number of moles present is:

$$n = PV/RT = (7.00 \times 10^4\ \text{Pa})(3.00 \times 10^{-3}\ \text{m}^3)/(8.314\ \text{m}^3\ \text{Pa K}^{-1}\ \text{mol}^{-1})(283.00\ \text{K})$$
$$= 8.925 \times 10^{-2}\ \text{mol}$$

At the new pressure, the volume occupied by this amount is:

$$V = nRT/P = (8.925\times10^{-2}\ \text{mol})(8.314\ \text{m}^3\ \text{Pa K}^{-1}\ \text{mol}^{-1})(283.00\ \text{K})/(1.01 \times 10^5\ \text{Pa})$$
$$= 2.08 \times 10^{-3}\ \text{m}^3 = 2.08\ \text{L}$$

More quickly, $P_1V_1 = P_2V_2$ can be used:

$$V_2 = P_1V_1 / P_2 = (7.00 \times 10^4\ \text{Pa})(3.00\ \text{L}) / (1.01 \times 10^5\ \text{Pa}) = 2.08\ \text{L}$$

Answer: **2.08 L**

Calculate the new pressure if the volume becomes 2.00 L and the temperature is unchanged.

From above, $n = 8.925 \times 10^{-2}$ mol. The pressure when $V = 2.00$ L and $T = 283.00$ K is:

$$P = nRT/V$$
$$= (8.925 \times 10^{-2}\ \text{mol})(8.314\ \text{m}^3\ \text{Pa K}^{-1}\ \text{mol}^{-1})(283.00\ \text{K})/(2.00 \times 10^{-3}\ \text{m}^3)$$
$$= 1.05 \times 10^5\ \text{Pa}$$

$P_1V_1 = P_2V_2$ can again be used without calculating n:

$$P_2 = P_1V_1 / V_2 = (7.00 \times 10^4\ \text{Pa}) \times (3.00\ \text{L}) / (2.00\ \text{L}) = 1.05 \times 10^5\ \text{Pa}$$

Answer: **1.05×10^5 Pa**

Calculate the new pressure if the temperature is raised to 50.0 °C and the volume is unchanged, *i.e.* still 3.00 L.

From above, $n = 8.925 \times 10^{-2}$ mol. The pressure when $V = 3.00$ L and $T = 323.0$ K is:

$$P = nRT/V$$
$$= (8.925 \times 10^{-2}\ \text{mol})(8.314\ \text{m}^3\ \text{Pa K}^{-1}\ \text{mol}^{-1})(323.0\ \text{K})/(3.00 \times 10^{-3}\ \text{m}^3)$$
$$= 7.99 \times 10^4\ \text{Pa}$$

The new pressure can be calculated directly using $P_1/T_1 = P_2/T_2$:

$$P_2 = P_1 \times T_2/T_1 = (7.00 \times 10^4\ \text{Pa}) \times 323.0/283.0 = 7.99 \times 10^4\ \text{Pa}$$

Answer: **7.99×10^4 Pa**

Marks
3

- A cylinder fitted with a piston contains 5.00 L of a gas at a pressure of 4.0×10^5 Pa. The entire apparatus is maintained at a constant temperature of 25 °C. The piston is released and the gas expands against a pressure of 1.0×10^5 Pa. Assuming ideal gas behaviour, calculate the final volume occupied by the gas.

As the number of moles and the temperature is constant, the initial and final pressures and volumes are related by:

$$\boldsymbol{V_1 P_1 = V_2 P_2}$$

Hence,

$$\boldsymbol{V_2 = V_1 P_1 / P_2 = (5.00\ \mathrm{L}) \times (4.0 \times 10^5\ \mathrm{Pa}) / (1.0 \times 10^5\ \mathrm{Pa}) = 20.\ \mathrm{L}}$$

Answer: **20. L**

Calculate the amount of work done by the gas expansion.

The gas expands from 5.00 to 20. L: it expands by 15 L. As 1 m^3 = 1000 L, this corresponds to 15×10^{-3} m^3.

The work done by a gas expanding against an external pressure is given by:

$$\boldsymbol{w = -P_{ext}\,\Delta V = -(1.0 \times 10^5\ \mathrm{Pa}) \times (15 \times 10^{-3}\ \mathrm{m^3}) = -1.5 \times 10^3\ \mathrm{J}}$$

Answer: **-1.5×10^3 J**

- The average speed of a gaseous neon atom at 300 K is 609 m s^{-1}. What is the average speed of a helium atom at the same temperature? **2**

As $E_{kinetic} = ½\, mv^2$:

$E_{kinetic}$ **(helium)** $= ½\, m_{He}v_{He}^2$

$E_{kinetic}$ **(neon)** $= ½\, m_{Ne}v_{Ne}^2$

The average kinetic energy of each gas is the same, at the same temperature, in the ideal gas model:

$½\, m_{He}v_{He}^2 = ½\, m_{Ne}v_{Ne}^2$

$v_{He}^2 = (m_{Ne} / m_{He}) \times v_{Ne}^2$

The ratio of the atomic masses is the same as the ratio of the molar masses and so:

$v_{He}^2 = (20.18 / 4.003) \times (609 \text{ m s}^{-1})^2$

$v_{He} = 1370 \text{ m s}^{-1}$

Answer: **1370 m s^{-1}**

- Why is helium instead of nitrogen mixed with oxygen in deep sea diving? Explain the origin of any differences in relevant properties. **2**

Increased pressure in lungs during deep sea diving causes an increased solubility of all gases in the blood. On ascending too quickly, these gases can bubble out of the blood. This is a serious problem with nitrogen as the bubbles can rupture blood vessels causing "the bends". The He atom is much smaller than the N_2 molecule, has a smaller electron cloud and is less polarisable. It therefore is less soluble in blood than nitrogen and is preferred as the above dangers are reduced.

Marks

- A doctor recommends to a pregnant woman that she takes an iron supplement of 50 mg (as Fe^{2+}) daily. To achieve this, what mass (to the nearest mg) of iron(II) gluconate-2-water, $FeC_{12}H_{22}O_{14}\cdot 2H_2O$, would be required? **3**

The atomic mass of Fe is 55.85 g mol^{-1}. a mass of 50 mg therefore corresponds to

$$\textbf{number of moles} = \frac{\textbf{mass}}{\textbf{atomic mass}} = \frac{50\times10^{-3}\ \text{g}}{55.85\ \text{g mol}^{-1}} = 8.95\times10^{4}\ \text{mol}$$

The molar mass of $FeC_{12}H_{22}O_{14}\cdot 2H_2O$ is:

molar mass = (55.85 (Fe) + 12×12.01 (C) + 26×1.008 (H) + 16×16.00 (O)) g mol^{-1}
= 482.178 g mol^{-1}

As 1 mole of this contains 1 mole of Fe, the mass of the supplement required is:

$$\textbf{mass} = \textbf{number of moles} \times \textbf{molar mass}$$
$$= (8.95\times10^{4}\ \text{mol})\times(482.178\ \text{g mol}^{-1}) = 0.432\ \text{g}$$

- What is the mass of each of the following at 298 K and 101 kPa pressure? **4**

(i) argon (24.5 litre)

Argon is a gas under these conditions. 24.5 L corresponds to the volume of 1.00 mol at 298 K and 101 kPa. Therefore, the mass of argon is:

mass = number of moles × atomic mass = (1.00 mol) × (39.95 g mol^{-1}) = 40.0 g

(ii) water (24.5 litre)

Water is a liquid under these conditions. Its density is 0.997 g cm^{-3}. The mass is therefore:

mass = density × volume = (0.997 g cm^{-3}) × (24.5 × 10^{3} cm^{3}) = 24400 g = 24.4 kg

(iii) chlorine (12.25 litre)

Cl_2 is a gas under these conditions. As 24.5 L corresponds to the volume of 1.00 mol, 12.25 L corresponds to $\dfrac{12.25\ \text{L}}{24.5\ \text{L mol}^{-1}} = 0.50\ \text{mol}$.

The molar mass of Cl_2 is (2 × 35.45 g mol^{-1}) = 70.9 g mol^{-1}. The mass is therefore:

mass = number of moles × atomic mass = (0.50 mol) × (70.9 g mol^{-1}) = 35.5 g

(iv) zinc (1.00 mole)

The atomic mass of Zn is 65.39 g mol^{-1}. The mass is therefore:

mass = number of moles × atomic mass = (1.00 mol) × (65.39 g mol^{-1}) = 65.4 g