Marks
4

- Henry's law relates the solubility of a gas to its pressure. *i.e.* c = *k*p

  The Henry’s law constant for N_2(g) at 298 K is 6.8×10^{-4} mol L^{-1} atm^{-1}. A diver descends to a depth where the pressure is 5 atm. If the diver’s body contains about 5 L of blood, calculate the maximum amount of nitrogen gas dissolved in the diver’s blood at 1 atm and at 5 atm. (Assume solubility of nitrogen in water and blood to be the same.)

At the surface, p = 1 atm and so:

$$c = kp = (6.8 \times 10^{-4} \text{ mol L}^{-1} \text{ atm}^{-1}) \times (1 \text{ atm}) = 6.8 \times 10^{-4} \text{ mol L}^{-1}$$

Hence, the amount of N_2 in 5 L is

$$\text{number of moles} = \text{concentration} \times \text{volume}$$
$$= (6.8 \times 10^{-4} \text{ mol L}^{-1}) \times (5 \text{ L}) = 0.003 \text{ mol}$$

When p = 5 atm and so:

$$c = (6.8 \times 10^{-4} \text{ mol L}^{-1} \text{ atm}^{-1}) \times (5 \text{ atm}) = 3.4 \times 10^{-3} \text{ mol L}^{-1}$$

Hence, the amount of N_2 in 5 L is

$$\text{number of moles} = (3.4 \times 10^{-4} \text{ mol L}^{-1}) \times (5 \text{ L}) = 0.02 \text{ mol}$$

1 atm: **3×10^{-3} mol**	5 atm: **2×10^{-2} mol**

If all the gas dissolved at 5 atm were suddenly released, what volume would it occupy at 1 atm and 298 K?

Using the ideal gas law, *PV* = *nRT*, with *n* = 2×10^{-2} mol:

$$V = nRT / P = (2 \times 10^{-2} \text{ mol})(0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1})(298 \text{ K}) / (1 \text{ atm}) = 0.4 \text{ L}$$

Answer: **0.4 L**