22 Marking scheme: Worksheet (A2)

1 **a** number of atoms = number of moles × N_A
number of atoms = $1.0 \times 6.02 \times 10^{23} \approx 6.0 \times 10^{23}$ [1]
b Number of molecules = $3.6 \times 6.02 \times 10^{23} \approx 2.2 \times 10^{24}$ [1]
c Number of atoms = $0.26 \times 6.02 \times 10^{23} \approx 1.6 \times 10^{23}$ [1]

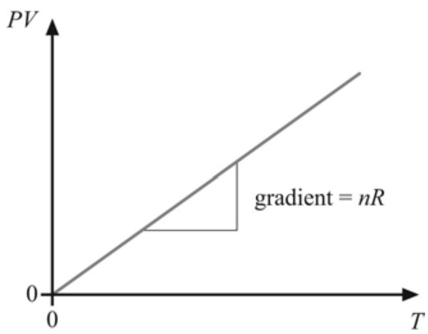
2 There are 6.02×10^{23} atoms in 4.0 g of helium. [1]
mass of atom = $\dfrac{0.004}{6.02 \times 10^{23}} = 6.645 \times 10^{-27}$ kg $\approx 6.6 \times 10^{-27}$ kg [1]

3 **a** There are 6.02×10^{23} atoms in 0.238 kg of uranium. [1]
mass of atom = $\dfrac{0.238}{6.02 \times 10^{23}} = 3.95 \times 10^{-25}$ kg $\approx 4.0 \times 10^{-25}$ kg [1]
b **i** number of moles = $\dfrac{\text{mass of uranium}}{\text{molar mass of uranium}}$ [1]
number of moles = $\dfrac{0.12}{238} = 5.04 \times 10^{-4} \approx 5.0 \times 10^{-4}$ [1]
ii number of atoms = number of moles × N_A
number of atoms = $5.04 \times 10^{-4} \times 6.02 \times 10^{23} = 3.06 \times 10^{20} \approx 3.1 \times 10^{20}$ [1]

4 The absolute zero of temperature is –273.15 °C or 0 K. [1]
This is the lowest temperature any substance can have. [1]
At absolute zero of temperature, the substance has minimum internal energy. [1]

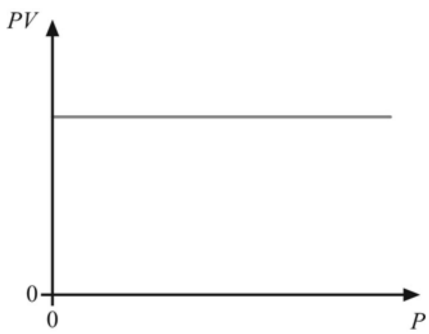
5 **a** Pressure × volume
= number of moles × universal gas constant × thermodynamic temperature [1]
b $PV = nRT$ [1]
$P = \dfrac{nRT}{V} = \dfrac{1.0 \times 8.31 \times 293}{0.020}$ [1]
$P = 1.22 \times 10^5$ Pa $\approx 1.2 \times 10^5$ Pa (120 kPa) [1]

6 **a** $PV = nRT$ [1]
Comparing this equation with $y = mx$, we have:
$y = PV$, $x = T$, gradient, $m = nR$ [1]
A graph of PV against T is a straight line through the origin.



Correct graph [1]
$n = \dfrac{\text{gradient}}{R}$ [1]

b $PV = nRT$ [1]
At a constant temperature, the product PV is a constant. [1]
Hence a graph of PV against P is a straight **horizontal** line. [1]



7 a $PV = nRT$ [1]

$n = \dfrac{4.0}{29} = 0.138$ moles [1]

$P = \dfrac{nRT}{V} = \dfrac{0.138 \times 8.31 \times (273 + 34)}{0.030}$ [1]

$P = 1.17 \times 10^4$ Pa ≈ 1.2×10^4 Pa (12 kPa) [1]

b $\dfrac{P}{T}$ is constant when the volume of the gas is constant. [1]

The pressure is doubled, hence the absolute temperature of the gas is also doubled. [1]
Therefore:
temperature = 2 × (273 + 34) = 614 K [1]
temperature in °C = 614 – 273 = 341 °C ≈ 340 °C [1]

8 a $n = \dfrac{PV}{RT}$ [1]

$n = \dfrac{180 \times 10^3 \times 2.0 \times 10^{-2}}{8.31 \times (273 - 13)} + \dfrac{300 \times 10^3 \times 2.0 \times 10^{-2}}{8.31 \times (273 - 13)}$ [1]

$n = 4.44$ moles ≈ 4.4 moles [1]

b Total volume, $V = 4.0 \times 10^{-2}$ m^3, $T = 273 - 13 = 260$ K

$P = \dfrac{nRT}{V}$ [1]

$P = \dfrac{4.44 \times 8.31 \times 260}{4.0 \times 10^{-2}}$ [1]

$P \approx 2.4 \times 10^5$ Pa (240 kPa) [1]

9 a $P = \dfrac{F}{A} = \dfrac{400}{1.6 \times 10^{-3}}$ [1]

$P = 2.5 \times 10^5$ Pa [1]

b $n = \dfrac{PV}{RT}$ [1]

$n = \dfrac{2.5 \times 10^5 \times 2.4 \times 10^{-4}}{8.31 \times (273 + 5.0)}$ [1]

$n = 2.6 \times 10^{-2}$ moles [1]

c i mass = number of moles × molar mass
mass = $2.6 \times 10^{-2} \times 29 = 0.754$ g ≈ 0.75 g [1]

ii density = $\dfrac{\text{mass}}{\text{volume}}$

density = $\dfrac{0.754 \times 10^{-3}}{2.4 \times 10^{-4}}$ [1]

density = 3.14 kg m^{-3} ≈ 3.1 kg m^{-3} [1]

10 Mean kinetic energy of atom $\propto$ absolute temperature [1]

$\frac{1}{2}mv^2 \propto T$ or $v^2 \propto \frac{2T}{m}$ [1]

Since the mass m of the atom is constant, we have: $v \propto \sqrt{T}$ [1]

The temperature of 0 °C in kelvin is $T = 273$ K

The absolute temperature increases by a factor of $\frac{10\,000}{273}$ (= 36.6) [1]

Hence the speed will increase by a factor of $\sqrt{\frac{10\,000}{273}} = 6.05$ [1]

The speed of the atoms at 10 000 K = 1.3 × 6.05 ≈ 7.9 km s^{-1} [1]

11 a The particles have a range of speeds and travel in different directions. [1]

b i Mean kinetic energy = $\frac{3}{2}kT = \frac{3}{2} \times 1.38 \times 10^{-23} \times 5400$ [1]

= 1.118 × 10^{-19} J ≈ 1.1 × 10^{-19} J [1]

ii $\frac{1}{2}mv^2 = \frac{3}{2}kT$ [1]

$v = \sqrt{\frac{3kT}{m}} = \sqrt{\frac{3 \times 1.38 \times 10^{-23} \times 5400}{1.7 \times 10^{-27}}}$ [1]

speed = 1.147 × 10^{4} m s^{-1} ≈ 11 km s^{-1} [1]

12 a Mean kinetic energy = $\frac{3}{2}kT = \frac{3}{2} \times 1.38 \times 10^{-23} \times 273$ [1]

= 5.65 × 10^{-21} J ≈ 5.7 × 10^{-21} J [1]

b There are 6.02 × 10^{23} molecules of carbon dioxide. [1]

mass of molecule = $\frac{0.044}{6.02 \times 10^{23}} = 7.31 \times 10^{-26}$ kg [1]

$\frac{1}{2}mv^2 = 5.65 \times 10^{-21}$ J [1]

$v = \sqrt{\frac{2 \times 5.65 \times 10^{-21}}{7.31 \times 10^{-26}}}$ [1]

speed = 393 m s^{-1} ≈ 390 m s^{-1} [1]

c Total kinetic energy of one mole of gas = $\frac{3}{2}kT \times N_A = \frac{3}{2}RT$ (Note: $R = k \times N_A$) [1]

For an ideal gas, the change in internal energy is entirely kinetic energy.

Change in internal energy = $\frac{3}{2}R \times (373 - 273) = \frac{3}{2} \times 8.31 \times 100$ [1]

change in internal energy = 1.2465 kJ ≈ 1.2 kJ [1]

13 a i The molecule has 3 degrees of freedom for translational motion and 2 degrees of freedom for rotation – making a total of 5. [1]

Therefore, mean energy = 5 × $\frac{1}{2}kT = \frac{5}{2}kT$ [1]

ii The molecule has 3 degrees of freedom for translational motion and 3 degrees of freedom for rotation – making a total of 6. [1]

Therefore, mean energy = 6 × $\frac{1}{2}kT = 3kT$ [1]

b Internal energy = $3kT \times N_A = 3RT$ (Note: $R = k \times N_A$) [1]

internal energy per unit kelvin = $3R$ [1]

= 3 × 8.31 ≈ 25 J K^{-1} [1]