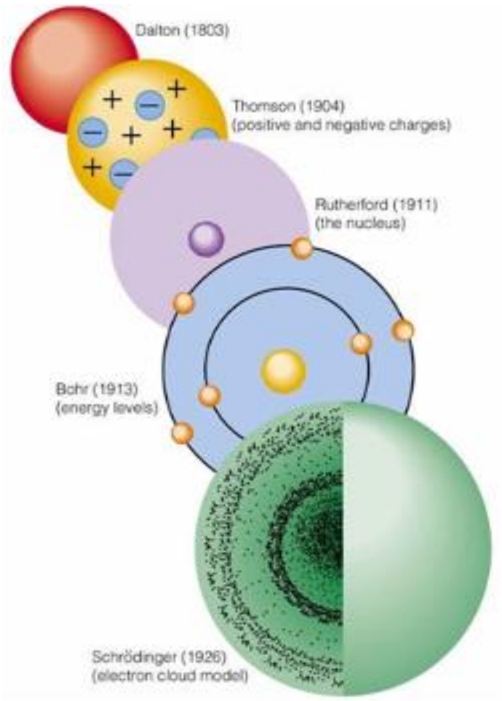
A level Chemistry Transition Booklet

Section 1: Atomic Structure

Our knowledge of the atom has changed over the last 200 years. Through investigations and new scientific technology, we have discovered more details about the atom.



Dalton: Solid sphere

Thompson: plum-pudding model (positive and negative charges.

Rutherford: Positive nucleus with electrons around it.

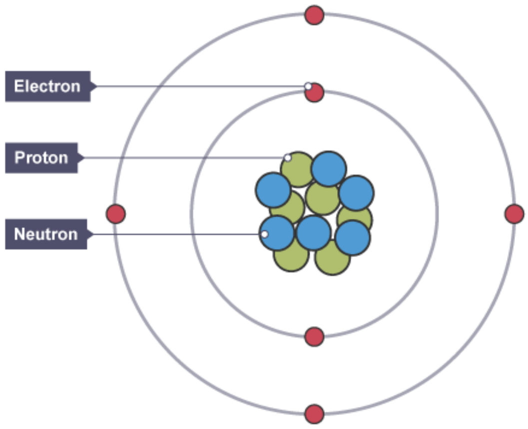
Bohr: Electrons in energy levels

Schrodinger: electrons in clouds (probable areas rather than rings).

A timeline of discovery

Date	Discovery
1897	Electrons
1909 - 1911	Atoms have a nucleus with electrons around it
1913	Electrons occupy shells (energy levels)
1918	Protons
1932	Neutrons

This is the atomic model we are familiar with; protons and neutrons are found in the nucleus and electrons in shells.



Mass Number = Number of protons and neutrons

Atomic number = Number of protons

Note: the charge on an atom is zero, as the protons and electron charges cancel out. This means there are the same amount of protons as neutrons.

Subatomic particle	Relative mass	Relative charge
Proton	1	+1
Neutron	1	0
Electron	$\frac{1}{1835}$	-1

To calculate the numbers of **subatomic particles** in an atom use its atomic number and mass number:

- number of protons = atomic number
- number of electrons = atomic number
- number of neutrons = mass number - atomic number

Isotopes: Atoms of the same element with a different atomic mass. This means the atoms have the same number of protons and electrons but a different number of neutrons.

The relative atomic mass (Ar) of an element is the weighted average of the relative atomic masses of the isotopes in the element. This is why on the periodic table, some of the atomic masses are shown as decimals.

To work out relative atomic mass:

$$A_r = \frac{total\ mass\ of\ atoms}{total\ number\ of\ atoms} = \frac{(35 \times 75) + (37 \times 25)}{(75 + 25)}$$

$$A_r = \frac{2625 + 925}{100} = \frac{3550}{100}$$

$A_r = 35.5$ (to 1 decimal place)

Practice Questions — Application

Use the periodic table to help you answer Questions 1-3.

Q1 Find the relative atomic mass of the following elements:

a) Rubidium

b) Mercury

c) Zinc

Q2 Find the relative molecular mass of the following compounds:

a) NH_3

b) CO_2

c) $C_2H_4O_6N_2$

Q3 Find the relative formula mass of the following compounds:

a) $CaCl_2$

b) $MgSO_4$

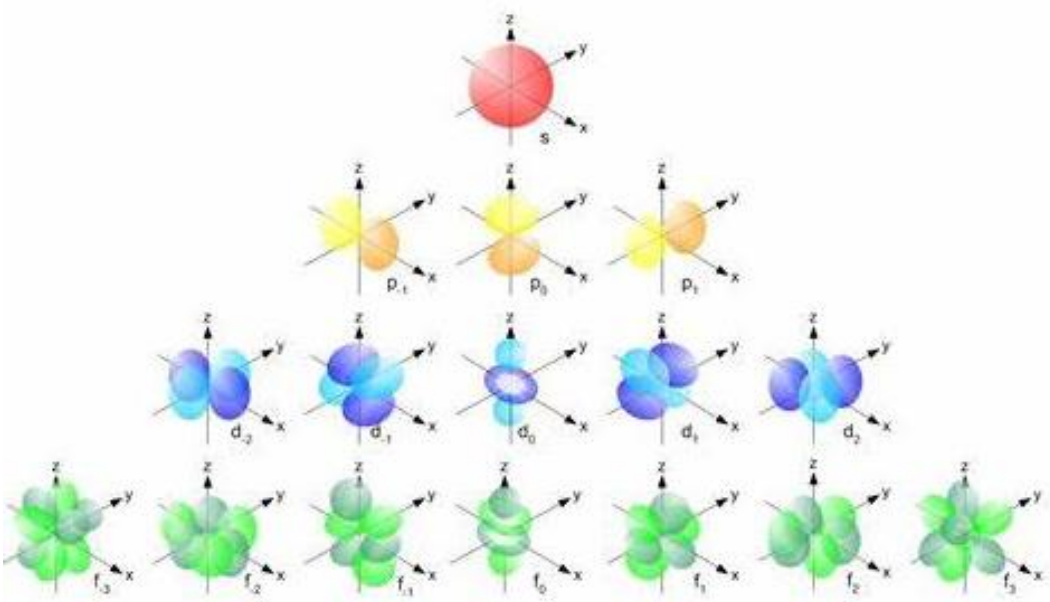
c) NaOH

Q4 A sample of tungsten is 0.1% ^{180}W, 26.5% ^{182}W, 14.3% ^{183}W, 30.7% ^{184}W and 28.4% ^{186}W. Calculate the A_r of tungsten.

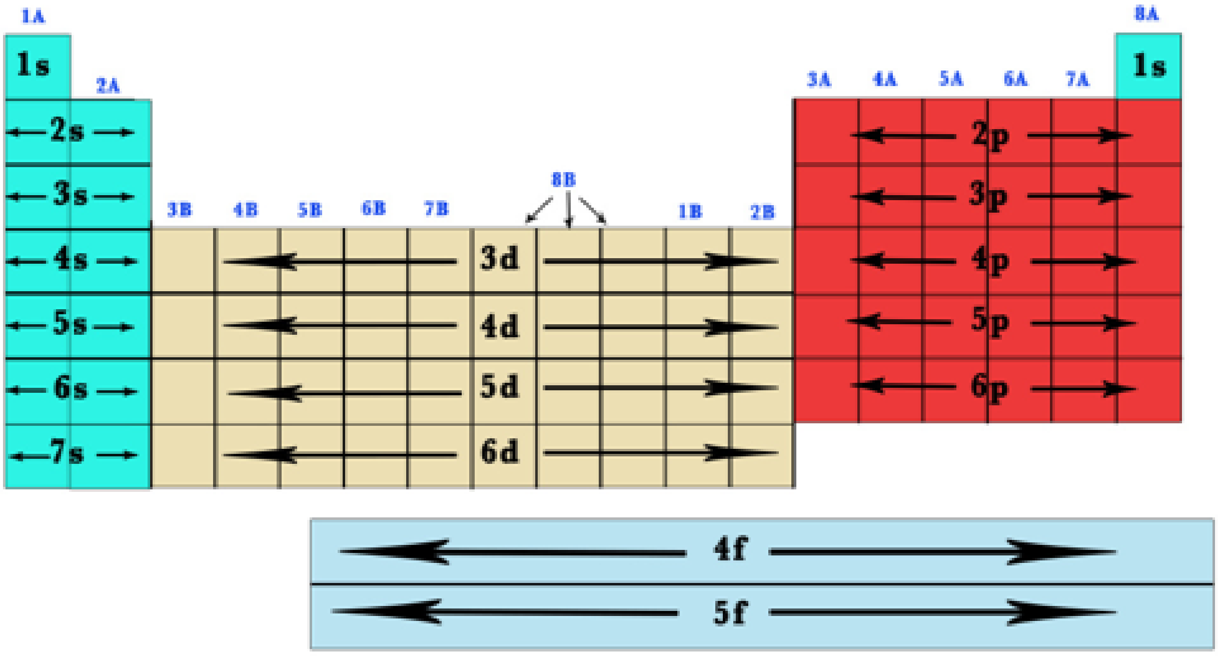
Q5 A sample of zirconium is 51.5% ^{90}Zr, 11.2% ^{91}Zr, 17.1% ^{92}Zr, 17.4% ^{94}Zr and 2.8% ^{96}Zr. Calculate the A_r of zirconium.

Electronic Structure

At GCSE, you learn the rule that electron shells can only hold a certain number of electrons. 1^{st} Shell holds 2, and the next three hold 8. However at a-level we learn that it is more complex than this. Each “shell” is made up of different sub-shells called s, p, d and f (see first diagram below). S shells are made of 1 orbital, p are made up of 3 orbitals, d are made up of 5 orbitals and f and made up of 7 orbitals. Each orbital can only hold a pair of electrons. Periodic table is split into the s, p, d and f blocks (see second diagram below).



s,p,d,f blocks in the periodic table.



So why do we teach you 2,8,8,8 at gcse?

Well we know that the period tells us the number of shells, so in the first period we can see there is only a 1s orbital. This means there is 2 electrons.

In the second period, we have 2s and 2p. The s orbital holds two electrons, and the p is made up of 3 orbitals, each holding 2 so 6 electrons. 6 + 2 is 8.

This is the same for the 3rd period, 3s and and 3p, so 2 + (3x2) = 8.

If we looked at the 4th period it is made up of 4s, 3d and 4p. 4s is one orbital so 2 electrons, 3d is 5 orbitals so 10 electrons and 4p is 3 orbitals which is 6. So the 4th “shell” holds 18 electrons.

Filling shells

Just like at GCSE, you must fill the lowest orbital before moving onto the next. The order is:1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p, 8s.

Note: you only need to know the order in red!

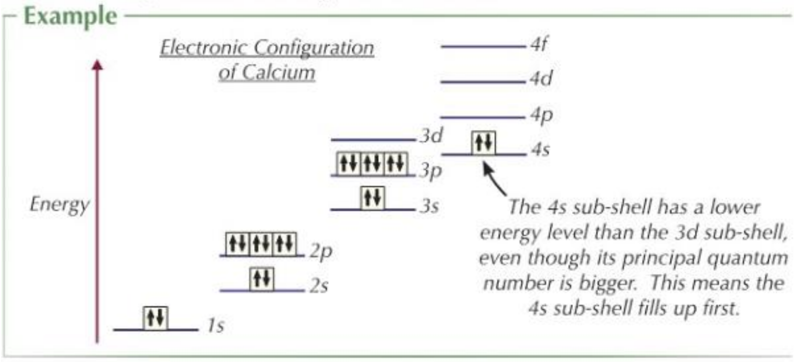
So let’s take an example, oxygen. Oxygen has 8 electrons. 1s will hold 2, 2s will hold 2 and 2p will hold 4. So we write 1s2, 2s2, 2p4. The first number and letter tell us the orbital, the last number (in red) tells us the number of electrons in that orbital.

You are not expected to draw the orbitals to show the electronic configuration, however you will be expected to show it in three ways:

1. Subshell notation: 1s2, 2s2, 2p6, 3s2, 3p4
2. Short-hand subshell notation: rather than writing the whole thing we can pick the nearest noble gas, in this case 1s2 2s2 2p6 = argon = [Ar]. We can then write the remaining subshells, [Ar] 3s2, 3p4.
3. Energy diagrams, (electrons in boxes): each box represents an orbital, lowest energy subshell at the bottom, highest energy subshell at the top. There are three rules for filling orbitals.

Rule 1

Electrons fill up the lowest energy sub-shells first.



Rule 2

Electrons fill orbitals in a sub-shell singly before they start sharing.

Rule 3

For the configuration of ions from the s and p blocks of the periodic table, just add or remove the electrons to or from the highest energy occupied sub-shell.

Examples

Mg atom: $1s^2\ 2s^2\ 2p^6\ 3s^2$

Mg^{2+} ion: $1s^2\ 2s^2\ 2p^6$

Cl atom: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

Cl^- ion: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Practice Questions — Application

Q1 Use sub-shell notation to show the full electron configurations of the elements listed below.

a) Lithium

b) Titanium

c) Gallium

d) Nitrogen

Q2 Draw arrows in boxes to show the electron configurations of the elements listed below.

a) Calcium

b) Nickel

c) Sodium

d) Oxygen

Q3 Draw energy level diagrams to show the electron configurations of the elements listed below.

a) Magnesium

b) Argon

c) Carbon

d) Arsenic

Q4 Use sub-shell notation to show the full electron configurations of the ions listed below.

a) Na^{+}

b) O^{2-}

c) Al^{3+}

d) S^{2-}

Q5 Which elements have the electron configurations given below?

a) $[Ar]3d^{10}\,4s^2\,4p^5$

b) $[Ne]3s^2\,3p^3$

c) $[Ar]3d^3\,4s^2$

Section 2: Amount of Substance

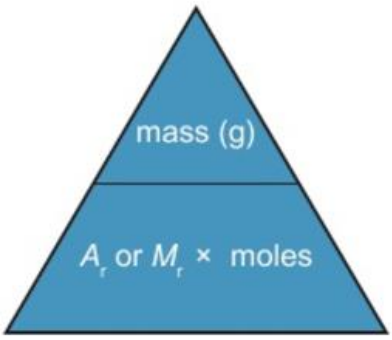
Moles

A mole of a substance contains 602 204 500 000 000 000 000 000 particles. This number is shortened to 6.02 X 1023 also known as Avagadro's Constant. The SI unit is mol.

The mass of one mole of a substance is its atomic mass (which we find in the periodic table). For example magnesium has an atomic mass of 24, therefore 1 mole of magnesium weighs 24 grams.

You can calculate the number of moles of a substance using the equation:

$$\text{number of moles of substance} = \frac{\text{mass of substance (g)}}{A_r \text{ or } M_r}$$



B To rearrange the equation with this triangle, cover up the quantity you want to calculate and what you can see gives you the calculation to use.

Number of particles = Number of moles × Avogadro's constant

Practice Questions — Application

Q1 How many molecules are in 0.360 moles of H_2O?

Q2 How many ions are in 0.0550 moles of magnesium ions?

Q3 1.5 moles of a mystery compound weighs 66 g. Find its relative molecular mass.

Q4 How many moles of sodium nitrate are present in 212.5 g of $NaNO_3$?

Q5 How many moles of zinc chloride are present in 15.5 g of $ZnCl_2$?

Q6 What is the mass of 2 moles of NaCl?

Moles and Concentration

$$\text{Number of moles} = \frac{\text{Concentration} \times \text{Volume (in cm}^3\text{)}}{1000}$$

Or: Number of moles = Concentration × Volume (in dm^3)

Practice Questions — Application

Q1 How many moles of potassium phosphate are present in 50 cm^3 of a 2 mol dm^{-3} solution?

Q2 How many moles of sodium chloride are present in 0.5 dm^3 of a 0.08 mol dm^{-3} solution?

Q3 How many moles of silver nitrate are present in 30 cm^3 of a 0.70 mol dm^{-3} solution?

Q4 A solution contains 0.25 moles of copper bromide in 0.50 dm^3. What is the concentration of the solution?

Q5 A solution contains 0.080 moles of lithium chloride in 0.75 dm^3. What is the concentration of the solution?

Q6 A solution contains 0.10 moles of magnesium sulfate in 36 cm^3. What is the concentration of the solution?

Q7 A solution of calcium chloride contains 0.46 moles of $CaCl_2$. The concentration of the solution is 1.8 mol dm^{-3}. What volume, in dm^3, does the solution occupy?

Q8 A solution of copper sulfate contains 0.010 moles of $CuSO_4$. The concentration of the solution is 0.55 mol dm^{-3}. What volume, in dm^3, does the solution occupy?

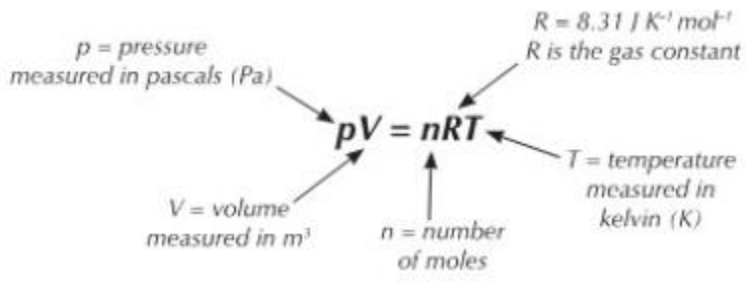
Q9 The molecular formula of sodium oxide is Na_2O. What mass of sodium oxide would you have to dissolve in 75 cm^3 of water to make a 0.80 mol dm^{-3} solution?

Q10 The molecular formula of cobalt(II) bromide is $CoBr_2$. What mass of cobalt(II) bromide would you have to dissolve in 30 cm^3 of water to make a 0.50 mol dm^{-3} solution?

Q11 A solution is made by dissolving 4.08 g of a compound in 100 cm^3 of pure water. The solution has a concentration of 1.20 mol dm^{-3}. What is the relative molecular mass of the compound?

Gases and the mole

Ideal Gas Equation



You could be asked to rearrange this in any way!

Unit conversions

Pressure (KPa to Pa) = multiply by 1000 e.g. 1KPa = 1000Pa

Temperature (Celcius to Kelvin) = add 273 e.g. 20C = 293K

Volume = 1m^3 is 1X10^6 cm^3 and 1m^3 is 1X10^3 dm^3

Practice Questions — Application

Q1 How many moles are there in 0.040 m^3 of oxygen gas at a temperature of 350 K and a pressure of 70 000 Pa?

Q2 What volume would 0.65 moles of carbon dioxide gas occupy at a temperature of 280 K and a pressure of 100 000 Pa?

Q3 How many moles are there in 0.55 dm^3 of nitrogen gas at a temperature of 35 °C and a pressure of 90 000 Pa?

Q4 At a pressure of 110 000 Pa, 0.0500 moles of hydrogen gas occupied a volume of 1200 cm^3. What was the temperature in °C?

Q5 What volume, in m^3, would 0.75 moles of helium gas occupy at a temperature of 22 °C and a pressure of 75 kPa?

Q6 At a temperature of 300 K and a pressure of 80 kPa a gas had a volume of 1.5 dm^3 and a mass of 2.6 g. Find its relative molecular mass.

Q7 A student had a sample of neon gas, Ne. They heated it to 44 °C. At this temperature the gas had a volume of 0.00300 m^3. If the pressure was 100 kPa, what was the mass of the neon gas?

Answers

Atomic Structure

Q1 a) 85.5
b) 200.6
c) 65.4

Q2 a) $14.0 + (3 \times 1.0) = \mathbf{17.0}$
b) $12.0 + (16.0 \times 2) = \mathbf{44.0}$
c) $(12.0 \times 2) + (1.0 \times 4) + (16.0 \times 6) + (14.0 \times 2) = \mathbf{152.0}$

Q3 a) $40.1 + (35.5 \times 2) = \mathbf{111.1}$
b) $24.3 + 32.1 + (16.0 \times 4) = \mathbf{120.4}$
c) $23.0 + 16.0 + 1.0 = \mathbf{40.0}$

Q4 $A_r = ((0.1 \times 180) + (26.5 \times 182) + (14.3 \times 183) + (30.7 \times 184) + (28.4 \times 186)) \div 100 = \mathbf{183.9}$ **(to 1 d.p.)**

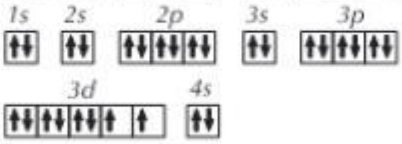
Q5 $A_r = ((51.5 \times 90) + (11.2 \times 91) + (17.1 \times 92) + (17.4 \times 94) + (2.8 \times 96)) \div 100 = \mathbf{91.3}$ **(to 1 d.p.)**

Electronic structure

Q1 a) $1s^2\ 2s^1$
b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$
(or $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$)
c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^1$
d) $1s^2\ 2s^2\ 2p^3$

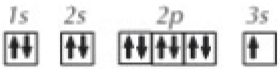
Q2 a)

b)

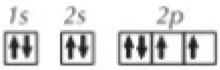


Remember that for each sub-shell you add, you should fill up each orbital singly before they start to share.

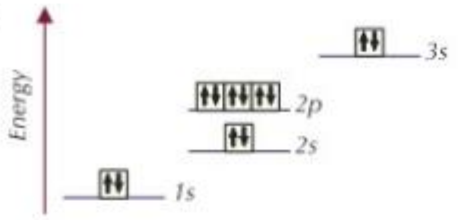
c)



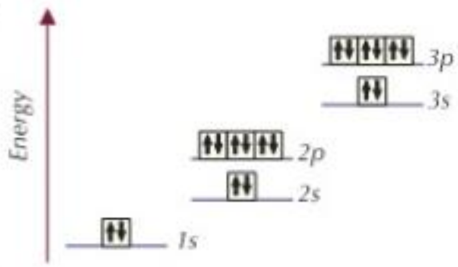
d)



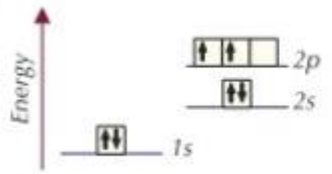
Q3 a)



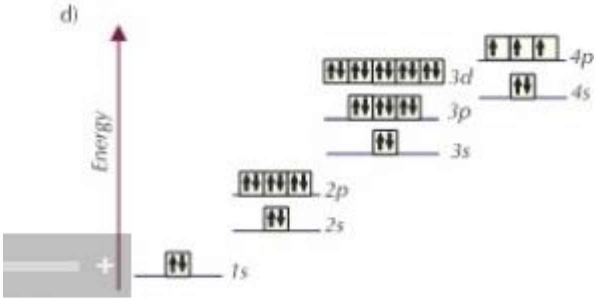
b)



c)



d)



Q4 a) $1s^2\ 2s^2\ 2p^6$
b) $1s^2\ 2s^2\ 2p^6$
c) $1s^2\ 2s^2\ 2p^6$
d) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Q5 a) bromine
b) phosphorus
c) vanadium

The mole

Q1 Number of molecules = $0.360 \times (6.02 \times 10^{23})$ = $\mathbf{2.17 \times 10^{23}}$
Q2 Number of ions = $0.0550 \times (6.02 \times 10^{23})$ = $\mathbf{3.31 \times 10^{22}}$
Q3 $M_r = 66 \div 1.5 = \mathbf{44}$
Q4 $M_r = 23.0 + 14.0 + (16.0 \times 3) = 85.0$
number of moles = $212.5 \div 85.0$ = **2.50 moles**
Q5 $M_r = 65.4 + (35.5 \times 2) = 136.4$
number of moles = $15.5 \div 136.4$ = **0.114 moles**
Q6 $M_r = 23.0 + 35.5 = 58.5$
Mass = 58.5×2 = **117 g**

Moles and concentration

Q1 Number of moles = $(2 \times 50) \div 1000$ = **0.1 moles**
Q2 Number of moles = 0.08×0.5 = **0.04 moles**
Q3 Number of moles = $(0.70 \times 30) \div 1000$ = **0.021 moles**
Q4 Concentration = $0.25 \div 0.50$ = $\mathbf{0.50\ mol\ dm^{-3}}$
Q5 Concentration = $0.080 \div 0.75$ = $\mathbf{0.11\ mol\ dm^{-3}}$
Q6 Concentration = $0.10 \div (36 \div 1000)$ = $\mathbf{2.8\ mol\ dm^{-3}}$

Q7 Volume = 0.46 ÷ 1.8 = **0.26 dm³**
Q8 Volume = 0.010 ÷ 0.55 = **0.018 dm³**
Q9 Number of moles = concentration × volume (dm³)
= 0.80 × (75 ÷ 1000) = 0.060
M_r of Na_2O = (23.0 × 2) + 16.0 = 62.0
Mass = moles × molar mass = 0.060 × 62.0 = **3.7 g**
Q10 Number of moles = concentration × volume (dm³)
= 0.50 × (30 ÷ 1000) = 0.015
M_r of $CoBr_2$ = 58.9 + (79.9 × 2) = 218.7
Mass = number of moles × molar mass
= 0.015 × 218.7 = **3.3 g**
Q11 Number of moles = concentration × volume (dm³)
= 1.20 × (100 ÷ 1000) = 0.120
M_r = mass ÷ number of moles
= 4.08 ÷ 0.120 = **34.0**

Gases and the Mole

Q1 $n = pV \div RT$
= (70 000 × 0.040) ÷ (8.31 × 350) = **0.96 moles**
Q2 $V = nRT \div p$
= (0.65 × 8.31 × 280) ÷ 100 000 = **0.015 m³**
Q3 0.55 dm³ = 5.5 × 10^{-4} m³ 35 °C = 308 K
$n = pV \div RT$
= (90 000 × (5.5 × 10^{-4})) ÷ (8.31 × 308)
= **0.019 moles**
Q4 1200 cm³ = 1.2 × 10^{-3} m³
$T = pV \div nR$
= (110 000 × (1.2 × 10^{-3})) ÷ (0.0500 × 8.31) = 318 K
318 K = (318 – 273) °C = **45 °C**
Q5 75 kPa = 75 000 Pa 22 °C = 295 K
$V = nRT \div p$
= (0.75 × 8.31 × 295) ÷ 75 000 = **0.025 m³**
Q6 80 kPa = 80 000 Pa 1.5 dm³ = 1.5 × 10^{-3} m³
$n = pV \div RT$
= (80 000 × (1.5 × 10^{-3})) ÷ (8.31 × 300)
= 0.048... moles
M_r = mass ÷ moles = 2.6 ÷ 0.048... = 54
So the relative molecular mass is **54**.
Q7 44 °C = 317 K 100 kPa = 100 000 Pa
$n = pV \div RT$
= (100 000 × 0.00300) ÷ (8.31 × 317)
= 0.113... moles
M_r of neon = 20.2
mass = number of moles × M_r
= 0.113... × 20.2 = **2.30 g**