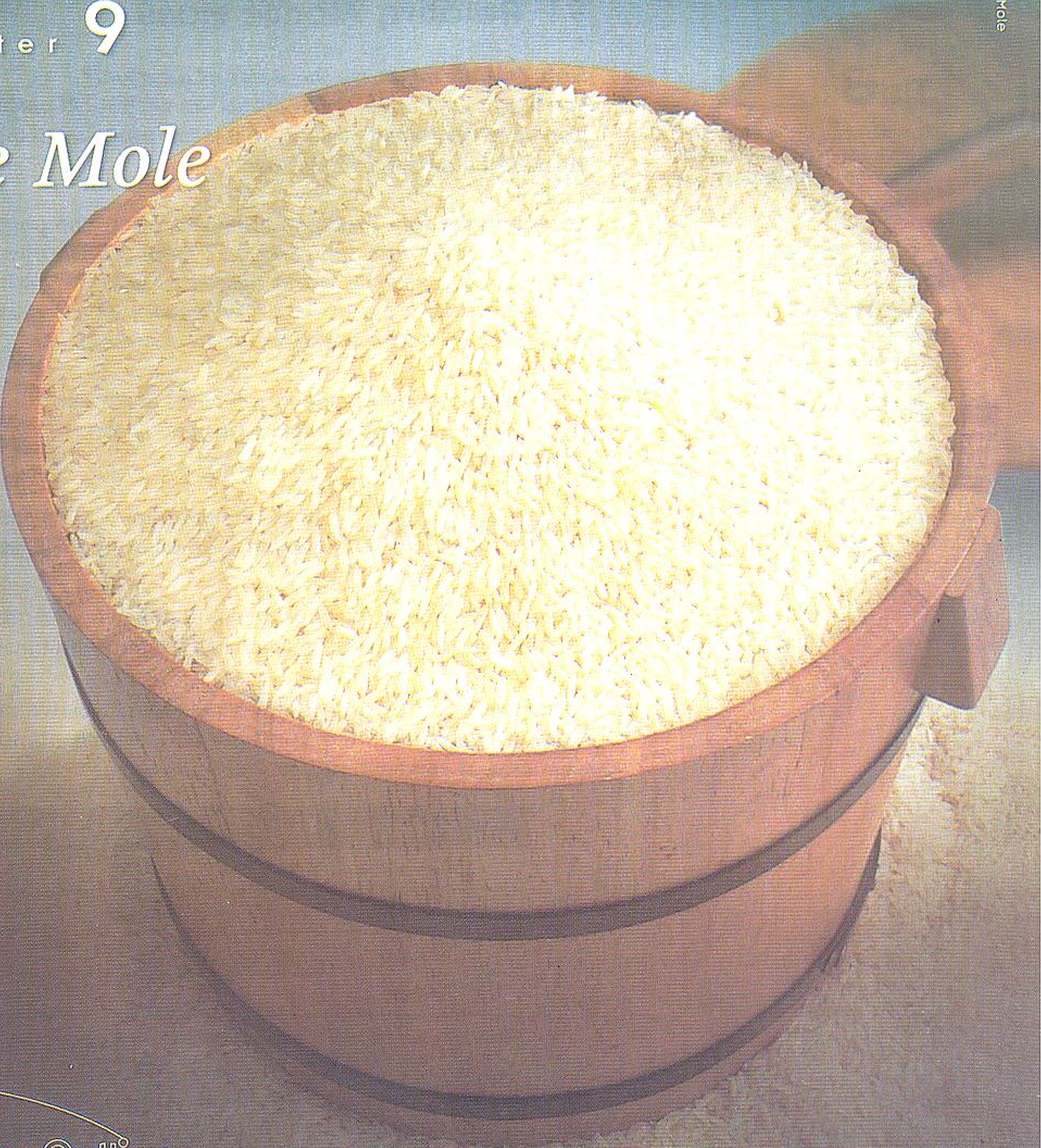
Chapter 9

The Mole



Chapter Outline

Have you ever tried counting the number of rice particles in a bucket of rice? It is difficult to do so because rice particles are very small and numerous. Chemists face a similar problem when they try to count atoms. Atoms are too small to be counted one at a time. Because they are so small, it is also difficult to measure the mass of each atom. In this chapter, you will find out how chemists overcome the problems of counting atoms and measuring the mass of each atom.

9.1 | Relative Atomic Mass

TidBit

The mass of a single carbon atom is about 0.000 000 000 000 000 000 000 000 020 04 g or 2.004×10^{-29} kg!

Chem-Aid

The carbon-12 atom has six protons, six neutrons and six electrons.

How can we measure the mass of an atom?

You learnt in chapter 4 that atoms are very small particles. Atoms also have very small masses, so it is not practical to use the actual masses of atoms in calculations.

To overcome this problem, chemists often compare masses of different atoms with the carbon-12 atom (an isotope of carbon). Scientists all over the world agreed to give the carbon-12 atom a relative atomic mass of 12. The masses of all other atoms are compared with one-twelfth the mass of one carbon-12 atom.

The **relative atomic mass** of any atom is *the number of times the mass of one atom of an element is greater than $\frac{1}{12}$ of the mass of one carbon-12 atom.*

$$\text{Relative atomic mass} = \frac{\text{mass of one atom of the element}}{\text{mass of } \frac{1}{12} \text{ of an atom of carbon-12}}$$

For example, one atom of oxygen is 16 times heavier than $\frac{1}{12}$ of an atom of carbon-12. Thus, oxygen has a relative atomic mass of 16.

The symbol for relative atomic mass is $\mathbf{A_r}$. Relative atomic mass is a ratio and therefore has *no unit*. The relative atomic masses of elements are given in the Periodic Table.

Element	Relative atomic mass, A_r
hydrogen	1
carbon	12
oxygen	16
chlorine	35.5

Table 9.1 Relative atomic masses of some common elements

Why are some A_r values not whole numbers?

As you can see from Table 9.1, the relative atomic mass of an element is usually a whole number. However, the relative atomic masses of some elements, such as chlorine, are not whole numbers. This is because such elements occur as mixtures of isotopes.

Link

Recall what you have learnt about isotopes in chapter 5.

For example, chlorine exists in two isotopic forms: chlorine-35 and chlorine-37. A sample of chlorine is made up of 75% of chlorine-35 atoms and 25% of chlorine-37 atoms.

Hence, relative atomic mass of chlorine

$$= \left(\frac{75}{100} \times 35\right) + \left(\frac{25}{100} \times 37\right)$$
$$= 26.25 + 9.25$$
$$= 35.5$$

9.2 | Relative Molecular Mass

Many elements and compounds exist as molecules. For example, chlorine exists as molecules. Each molecule of chlorine consists of two chlorine atoms. One molecule of nitrogen dioxide consists of one nitrogen atom and two oxygen atoms (Fig. 9.1).

Fig. 9.1 Chlorine and nitrogen dioxide exist as molecules.

The mass of a molecule is measured in terms of its relative molecular mass. The **relative molecular mass (M_r)** of an element or compound is *the mass of a molecule*, compared to $\frac{1}{12}$ *the mass of one atom of carbon-12.*

Relative molecular mass (M_r)

$$= \frac{\text{mass of one molecule of an element or compound}}{\text{mass of } \frac{1}{12} \text{ of an atom of carbon-12}}$$

How do we calculate the relative molecular mass of a molecule?

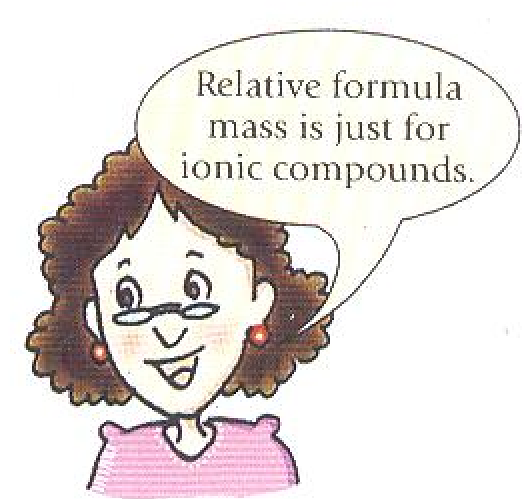
The relative molecular mass of a molecule is calculated by adding together the relative atomic masses of each atom in its chemical formula (Table 9.2). Like relative atomic mass, it is a ratio and therefore has *no unit.*

Molecule	Chemical formula	Number of atoms in one molecule	Calculating M_r
nitrogen	N_2	2 N	(2 × 14) = **28**
ammonia	NH_3	1 N; 3 H	(1 × 14) + (3 × 1) = **17**
carbon dioxide	CO_2	1 C; 2 O	(1 × 12) + (2 × 16) = **44**
water	H_2O	2 H; 1 O	(2 × 1) + (1 × 16) = **18**
ethanol	C_2H_5OH	2 C; 6 H; 1 O	(2 × 12) + (6 × 1) + (1 × 16) = **46**

Table 9.2 Calculating the relative molecular masses of some molecules

Relative Formula Mass

You learnt in chapter 7 that substances like water are covalent and exist as molecules. However, substances like sodium chloride are ionic and do not exist as molecules.



The relative molecular mass of an ionic compound is more accurately known as the **relative formula mass**. Like relative molecular mass, relative formula mass is given the symbol M_r and has *no units*. For example, the relative formula mass of sodium chloride (NaCl) is 23 + 35.5 = 58.5.

Substance	Formula unit	Number of atoms in formula unit	Calculating M_r
magnesium sulphate	$MgSO_4$	1 Mg; 1 S; 4 O	$(1 \times 24) + (1 \times 32) + (4 \times 16) =$ **120**
calcium carbonate	$CaCO_3$	1 Ca; 1 C; 3 O	$(1 \times 40) + (1 \times 12) + (3 \times 16) =$ **100**
calcium nitrate	$Ca(NO_3)_2$	1 Ca; 2 N; 6 O	$(1 \times 40) + (2 \times 14) + (6 \times 16) =$ **164**
copper(II) sulphate crystals	$CuSO_4.5H_2O$	1 Cu; 1 S; 9 O; 10 H	$(1 \times 64) + (1 \times 32) + (9 \times 16) + (10 \times 1) =$ **250**

Table 9.3 Calculating the relative formula masses of some ionic substances

Chem-Aid

The dot ('.') in $CuSO_4.5H_2O$ means that there are five H_2O molecules bonded to each $CuSO_4$. It does not mean multiply.

Key ideas

1. The relative atomic mass (A_r) of any atom is the number of times the mass of one atom of an element is heavier than $\frac{1}{12}$ of a carbon-12 atom.
2. A_r values may not be whole numbers because of isotopes.
3. The relative molecular mass (M_r) of a molecule is the sum of the relative atomic masses of all the atoms in the molecule.
4. The relative formula mass (M_r) of an ionic compound is the sum of the relative atomic masses of atoms in a formula unit of the compound.

Test Yourself 2.1

Worked Example

$^{24}_{12}S$ and $^{16}_{8}T$ are two elements that react together to form an ionic compound *V*. What is the relative formula mass (M_r) of *V*?

A 20 B 28 C 32 D 40

Thought Process

The electronic configuration of *S* is (2, 8, 2). *S* has a valency of 2. The electronic configuration of *T* is (2, 6). *T* also has a valency of 2. Hence, the formula of this compound is *ST*. Accordingly, the M_r of *V* is 24 + 16 = 40.

Answer

D

Questions

1. Calculate the relative molecular mass/relative formula mass of each of the following.
   a) $Ca(OH)_2$
   b) $(NH_4)_3PO_4$
   c) $C_2H_5COOCH_3$
   d) $Pb(NO_3)_2$
   e) $Na_2CO_3.10H_2O$
2. Four nitrogen atoms have the same mass as one formula unit of *XO*. What is *X*?

9.3 | The Mole

Atoms are too small and numerous to be counted one at a time. Instead, the quantity of atoms is measured by mass. The unit of measurement for atoms and molecules is the **mole**. The mole is also the S.I. unit for chemical quantity. The symbol for the mole is **mol**.

A mole of substance contains the *same number of particles as the number of atoms in 12 g of carbon-12.*

How many particles are there in a mole?
There are *approximately* 6×10^{23} particles in one mole of substance. 6×10^{23} is called the **Avogadro's constant**. One mole of particles contains 6×10^{23} particles. The particles could be atoms, molecules, ions or electrons.

Chem-Aid

6×10^{23} is known as Avogadro's constant. It may also be known as Avogadro's number.

One mole of particles contains 6×10^{23} particles.

Even if you started counting now at the rate of 10 million atoms per second, it would take you 2 billion years to count up to Avogadro's number!

How do we convert between number of moles and number of particles?
Since one mole of substance contains 6×10^{23} particles,

$$\text{Number of moles} = \frac{\text{number of particles}}{6 \times 10^{23}}$$

Equal numbers of moles contain equal numbers of particles. The reverse is also true.

Example 1

Convert 1×10^{23} of neon atoms to moles of neon atoms.

Solution:

$$\begin{aligned}\text{Number of moles of neon atoms} &= \frac{\text{number of neon atoms}}{\text{Avogadro's constant}} \\ &= \frac{1 \times 10^{23}}{6 \times 10^{23}} \\ &= \mathbf{0.167\ mol}\end{aligned}$$

Example 2

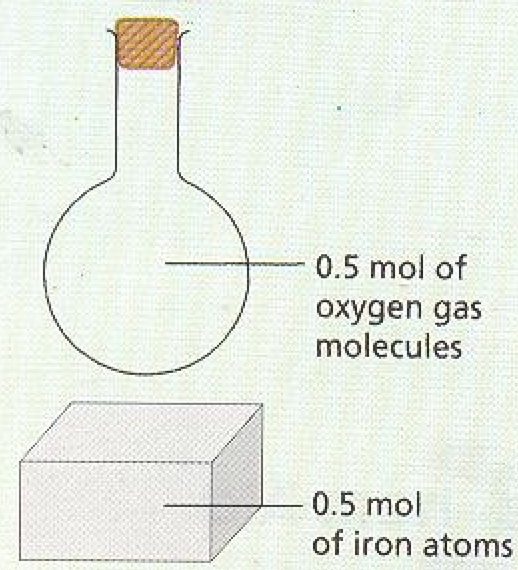
How many iron atoms are there in 0.5 mol of iron?

Solution:

In one mole of iron, there are 6×10^{23} iron atoms.

$$\begin{aligned}\text{Number of iron atoms} &= \text{number of moles} \times 6 \times 10^{23} \\ &= 0.5 \times 6 \times 10^{23} \\ &= \mathbf{3 \times 10^{23}}\end{aligned}$$

Quick Check



Are the number of particles in 0.5 mol of oxygen gas molecules and 0.5 mol of iron atoms the same?

Example 3

How many hydrogen atoms are there in three moles of hydrogen gas?

Solution:

Hydrogen gas is made up of hydrogen molecules (H_2).
In one mole of H_2 molecules, there are two moles of H atoms.
In three moles of H_2 molecules, there are six moles of H atoms.

Therefore, the number of hydrogen atoms
$= 6 \times 6 \times 10^{23}$
$= \mathbf{3.6 \times 10^{24}}$

How is the mole related to chemical equations?

Consider the following chemical equation:

$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

It tells us that: two particles (atoms) of sodium → react with → one particle (molecule) of chlorine → to form → two units of sodium chloride

OR

in terms of moles: two moles of sodium atoms → react with → one mole of chlorine molecules → to form → two moles of sodium chloride

Understanding that equal numbers of moles contain equal numbers of particles is important when we perform chemical calculations. You will learn more about this in the next chapter.

9.4 | Mole and Molar Mass

What is the mass of one mole of atoms of an element?

The **molar mass** of an element is the *mass of one mole* of atoms of the element. Look at Table 9.4. Do you notice a relationship between the value of A_r and the molar mass of a substance?

Element	A_r	Molar mass
aluminium	27	27 g
carbon	12	12 g
neon	20	20 g
oxygen	16	16 g

Table 9.4 The molar masses of some elements

Chem-Aid

The unit for molar mass may be 'g' or 'g/mol'.

The molar mass is equal to the relative atomic mass (A_r) of the element in grams. For example, the relative atomic mass of sodium is 23. The mass of one mole of sodium atoms is 23 g. We can also say that the molar mass of sodium is 23 g.

What is the relationship between mole and molar mass?

The number of moles of an element can be calculated using the formula:

$$\text{Number of moles of an element} = \frac{\text{mass of element in grams}}{\text{relative atomic mass of the element}}$$

Example 1

Determine the number of moles in 0.196 kg of iron. (A_r: Fe = 56)

Solution:

$$\text{Number of moles of iron} = \frac{\text{mass of iron in grams}}{A_r \text{ of iron}}$$

$$= \frac{0.196 \times 1000}{56}$$

$$= \mathbf{3.5\ mol}$$

Chem-Aid

Convert the unit of mass to grams. In this example, the mass is given in kg, therefore it has to be multiplied by 1000 to convert it to grams.

Example 2

a) How many moles of lead are there in 1.204×10^{22} atoms of lead?
b) What is the mass of 1.204×10^{22} atoms of lead? (A_r: Pb = 207)

Solution:

a) $$\text{Number of moles of lead} = \frac{\text{number of atoms of lead}}{\text{Avogadro's constant}}$$

$$= \frac{1.204 \times 10^{22}}{6 \times 10^{23}}$$

$$= \mathbf{0.02\ mol}$$

b) $$\text{Mass of lead} = \text{number of moles} \times A_r \text{ of lead}$$

$$= 0.02 \times 207$$

$$= \mathbf{4.14\ g}$$

Quick Check

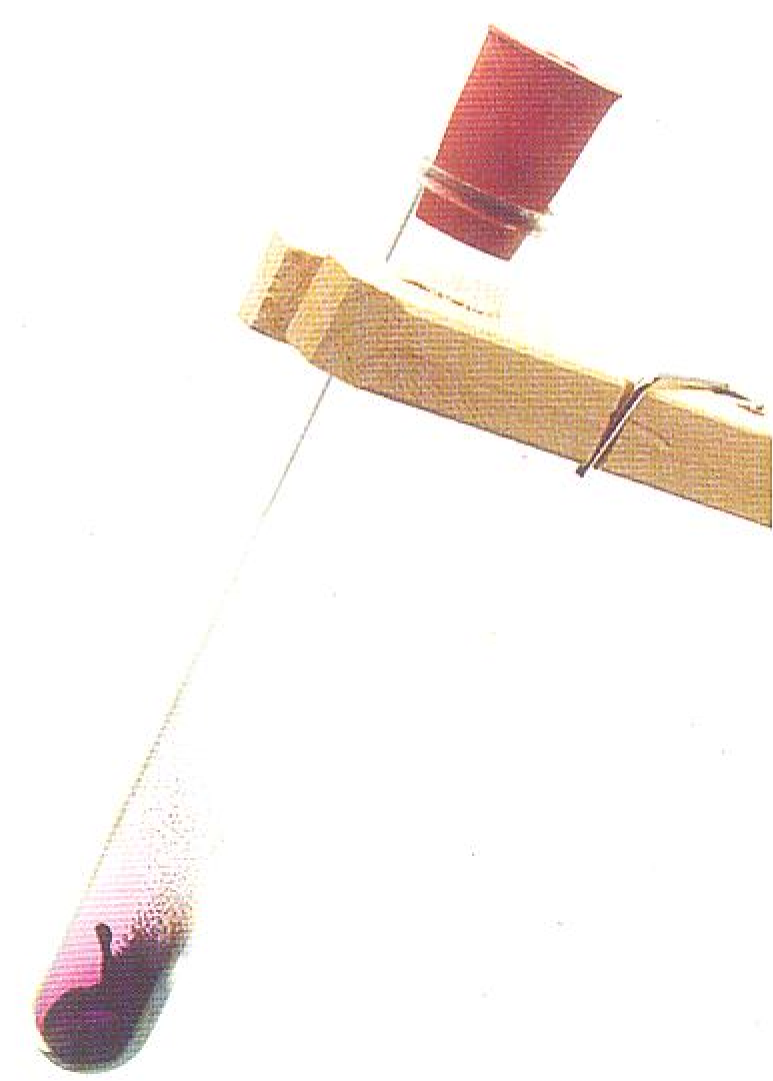
Which has the greater mass, 1.0 g of gold or 0.5 mol of helium? Show your working.

What is the mass of one mole of molecules or one mole of a compound?

In the previous section, you learnt that the mass of one mole of atoms is the same as its A_r in grams. The same idea can be extended to molecules and compounds. Table 9.5 shows the molar masses of some common elements and compounds. One mole of a substance will have a mass equal to the relative molecular mass or relative formula mass in grams.

Substance	Formula	M_r	Molar mass
oxygen	O_2	$2 \times 16 = 32$	32 g
iodine	I_2	$2 \times 127 = 160$	254 g
magnesium oxide	MgO	$24 + 16 = 40$	40 g
water	H_2O	$(2 \times 1) + 16 = 18$	18 g

Table 9.5 The molar masses of common substances



One mole of iodine (I_2) weighs 254 g.

The number of moles of a substance can be calculated using the formula:

$$\text{Number of moles} = \frac{\text{mass of substance (g)}}{\text{relative molecular mass}}$$

OR

$$= \frac{\text{mass of substance (g)}}{\text{relative formula mass}}$$

Remember, we have also learnt that

$$\text{Number of moles} = \frac{\text{number of particles}}{\text{Avogadro's constant}}$$

Example 1

A conical flask contains 68.4 g of octane (C_8H_{18}). How many molecules of octane are there in the flask?

Solution:

$$\text{Number of moles of octane} = \frac{\text{mass of octane}}{M_r \text{ of octane}}$$

$$= \frac{68.4}{114}$$

$$= \mathbf{0.6\ mol}$$

Number of molecules of octane

= number of moles × Avogadro's constant

= $0.6 \times 6 \times 10^{23}$

= $\mathbf{3.6 \times 10^{23}}$

Example 2

How many **ions** are there in 20 g of magnesium oxide (MgO)?

Solution:

Number of moles of magnesium oxide

$$= \frac{\text{mass of magnesium oxide}}{M_r \text{ of magnesium oxide}}$$

$$= \frac{20}{40}$$

$$= \mathbf{0.5\ mol}$$

$$MgO \longrightarrow Mg^{2+} + O^{2-}$$

From the equation, we can see that 1 mol of MgO contains 1 mol of Mg^{2+} ions and 1 mol of O^{2-} ions. 0.5 mol of MgO will contain 0.5 mol of Mg^{2+} ions and 0.5 mol of O^{2-} ions.

Hence, 0.5 mol of MgO contains 1 mol of ions.

Number of ions = $1 \times 6 \times 10^{23}$ = $\mathbf{6 \times 10^{23}}$

Key ideas

1. The mass of one mole of a substance is its
   a) relative atomic mass in grams if the substance exists as atoms e.g. neon (Ne).
   b) relative molecular mass in grams if the substance exists as molecules e.g. oxygen (O_2).
   c) relative formula mass in grams if the substance is an ionic compound e.g. magnesium oxide (MgO).
2. A mole of any substance contains 6×10^{23} particles. This number is called the Avogadro's constant or Avogadro's number.
3. Number of moles of atoms $= \dfrac{\text{mass of element (g)}}{A_r}$

   Number of moles of substance $= \dfrac{\text{mass of substance (g)}}{M_r}$
4. Molar mass refers to the mass of one mole of a substance. It has the same value as A_r or M_r.

Test Yourself 9.2

Worked Example

A metal compound has the formula XCl_4 and a relative formula mass of 261. What is

a) the relative atomic mass of *X*?

b) metal *X*?

Thought Process

a) Let *M* be the A_r of metal *X*.

M_r of $XCl_4 = M + 4 \times 35.5 = M + 142$

Therefore, $M + 142 = 261$

$M = 261 - 142 = 119$

b) Using the Periodic Table, the metal with an A_r of 119 is tin (Sn).

Answer

a) 119

b) Tin

Questions

1. A crucible contains 254 g of iodine I_2. How many moles of
   a) iodine molecules and
   b) iodine atoms are there in the crucible?

2. What is the mass of
   a) 0.25 mol of oxygen gas?
   b) 0.25 mol of nitrate ions (NO_3^-)?

3. Calculate
   a) the number of moles in 36 g of carbon.
   b) the mass of 0.4 mol of hydrogen sulphide (H_2S).
   c) the number of atoms in 6 g of magnesium.

9.5 | Percentage Composition of Compounds

Look at the pies in the picture. How can you tell which pie has more filling? You would need to examine each pie, for example, by breaking it into two. This is a form of analysis.

Chemists need to conduct analysis too, in order to find out how much of each element there is in a new compound. They do so by finding out the mass of each element in the compound. In this way, chemists know the percentage composition of a compound. Let us analyse the percentage composition of the compound hydrogen peroxide (H_2O_2).

How do we find the percentage composition of hydrogen peroxide?

In general, the percentage by mass of an element in a compound can be found using the formula:

Percentage by mass of an element in a compound

$$= \frac{A_r \text{ of element} \times \text{number of atoms in formula}}{\text{relative molecular mass } (M_r) \text{ of compound}} \times 100\%$$

The percentages (by mass) of hydrogen and oxygen present in hydrogen peroxide can be calculated from its chemical formula, H_2O_2.

Relative molecular mass of hydrogen peroxide (H_2O_2)

$= (2 \times 1) + (2 \times 16)$

$= 34$

Percentage of hydrogen in hydrogen peroxide

$$= \frac{A_r \text{ of hydrogen} \times \text{number of hydrogen atoms}}{\text{relative molecular mass of } H_2O_2} \times 100\%$$

$$= \frac{1 \times 2}{34} \times 100\%$$

$= \mathbf{5.9\%}$

Percentage of oxygen in hydrogen peroxide

$$= \frac{A_r \text{ of oxygen} \times \text{number of oxygen atoms}}{\text{relative molecular mass of } H_2O_2} \times 100\%$$

$$= \frac{16 \times 2}{34} \times 100\%$$

$= \mathbf{94.1\%}$

Example 1

Calculate the percentage of water in copper(II) sulphate crystals ($CuSO_4.5H_2O$).

Solution:

M_r of copper(II) sulphate crystals
$= 64 + 32 + (4 \times 16) + (5 \times 18)$
$= 250$

M_r of water
$= (1 \times 2) + 16$
$= 18$

$\therefore$ Percentage of water

$$= \frac{M_r \text{ of water} \times \text{number of water molecules}}{M_r \text{ of copper(II) sulphate crystals}} \times 100\%$$

$$= \frac{18 \times 5}{250} \times 100\%$$

$= \mathbf{36\%}$

Key ideas

1. The percentage composition of a compound can be found given
   a) its formula,
   b) the relative atomic masses of elements in it.
2. Percentage composition by mass of an element in a compound

$$= \frac{A_r \text{ of element} \times \text{number of atoms of element in formula}}{M_r \text{ of compound}} \times 100\%$$

Test Yourself 9.3

Question

Calculate the following:

a) The percentage of nitrogen in potassium nitrate, KNO_3.
b) The percentage of chlorine in ammonium chloride, NH_4Cl.
c) The percentages of calcium and oxygen in calcium carbonate, $CaCO_3$.

9.6 Finding the Formula of a Compound

Do you remember reading about the discovery of aniline in chapter 2? Chemists often discover useful new compounds while carrying out different tasks. In order to produce these useful products on a large scale, chemists need to know the chemical formulae of these compounds.

How can we work out the formula of a compound?

In the previous section, we saw that it was relatively easy to determine the percentage composition of a substance. This is because we already knew the formula of the compound. What if we did not know the formula of the compound?

We can conduct experiments to find out the formula of a compound.

1. First, we find out the mass of the reactants taking part in the reaction.
2. Next, we work out the relative numbers of moles of the reactants used.

We are then able to find the formula of the compound.

Experiment 1

To work out the formula of magnesium oxide produced by the combustion of magnesium

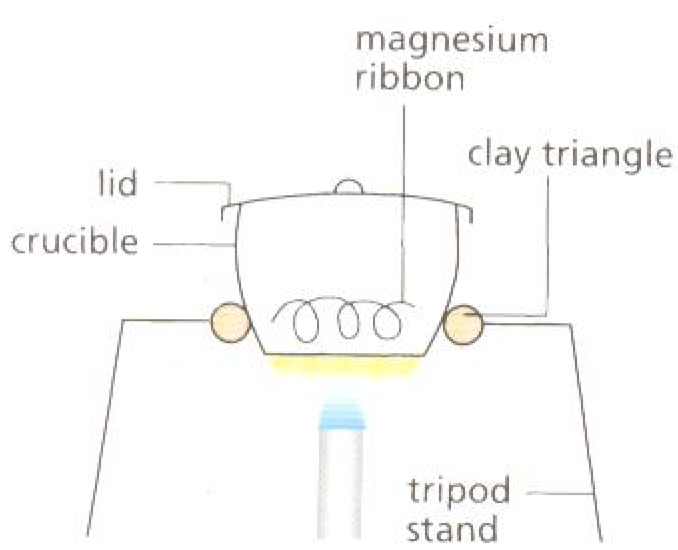


Fig. 9.2 Finding the formula of magnesium oxide

Procedure

1. Weigh a crucible together with the lid. Put a coil of magnesium ribbon in it and weigh again.
2. Put the lid on the crucible and heat the crucible gently (Fig. 9.2). When the magnesium catches fire (you will see a white glow through the crucible), heat it more strongly.
3. Use a pair of tongs to lift the lid slightly from time to time to allow air in. Quickly replace the lid to make sure that magnesium oxide formed does not escape.
4. When the burning is complete, allow the crucible to cool. Then weigh the crucible together with the lid and the magnesium oxide in it.

Sample results

Mass of crucible + lid = 26.52 g
Mass of crucible + lid + magnesium = 27.72 g
Mass of crucible + lid + magnesium oxide = 28.52 g

Calculations

Mass of magnesium = 27.72 − 26.52 = 1.20 g
Mass of magnesium oxide produced = 28.52 − 26.52 = 2.00 g
Mass of oxygen reacted = 2.00 − 1.20 = 0.80 g

Table 9.6 shows the masses of magnesium and oxygen that reacted in the experiment.

Element	Mg	O
Mass (obtained from experiment)	1.20 g	0.80 g
Relative atomic mass	24	16
Number of moles	$\frac{1.20}{24} = 0.05$	$\frac{0.80}{16} = 0.05$
Molar ratio (divide by smallest number from previous row)	$\frac{0.05}{0.05} = 1$	$\frac{0.05}{0.05} = 1$

smallest value in the 'number of moles' row

Table 9.6 Finding the number of moles of magnesium and oxygen that reacted

The simplest formula of magnesium oxide is thus **MgO**.

Empirical Formula

In experiment 1, we found that the formula MgO is the simplest formula that fits the experimental results. Other formulae such as Mg_2O_2 or Mg_3O_3 also fit the results since the ratios of magnesium to oxygen are the same. The formula MgO is called *the simplest formula or the empirical formula* of magnesium oxide. The **empirical formula** of a compound shows

- the types of elements present in it,
- the simplest ratio of the different types of atoms in it.

Chem-Aid

1. Remember:
   number of moles
   $= \frac{\text{mass}}{M_r}$

2. Notice that it is relative atomic mass, not relative molecular mass, that is used in deriving empirical formula.

Experiment 2

To determine the empirical formula of copper(II) oxide

Procedure

1. Weigh a porcelain boat. Put one spatula of copper(II) oxide (a black powder) in the boat and weigh again.
2. Put the porcelain boat containing copper(II) oxide into the middle of a Pyrex test tube with a small jet hole at its end.
3. Set up the apparatus that is shown in Fig. 9.3. Allow methane gas (from the gas tap) to pass through the apparatus for about 30 seconds to remove air.
4. Light the gas that escapes through the hole in the test tube. Then heat the copper(II) oxide until no further colour change is observed. The following reaction occurs during heating:

   copper(II) oxide + methane ⟶ copper + carbon dioxide + water

   Copper will remain as a brown residue in the porcelain boat.
5. Turn off the Bunsen burner but allow the methane gas to flow through the apparatus while the apparatus is cooling.
6. Turn off the gas supply when the apparatus is cool. Weigh the porcelain boat and its contents.

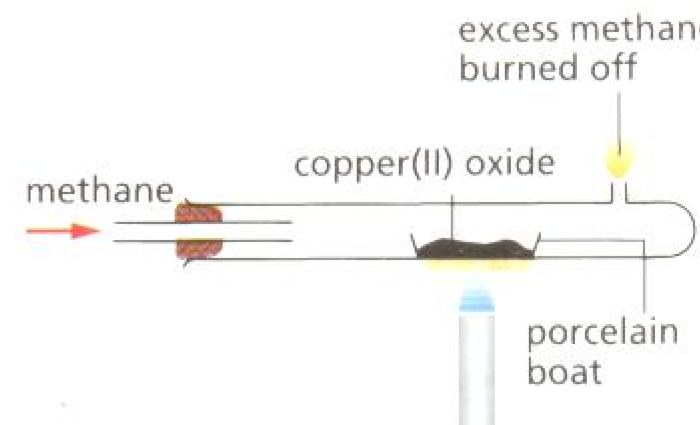


Fig. 9.3 Finding the formula of copper(II) oxide

Science Skills

a) Why is it necessary to remove all the air from the apparatus in Fig. 9.3?

b) Why is it necessary to allow the methane gas to flow while the apparatus is cooling?

Sample results

Mass of porcelain boat = 18.40 g
Mass of porcelain boat + copper(II) oxide = 35.89 g
Mass of porcelain boat + copper = 32.37 g

Calculations

Mass of copper = 32.37 − 18.40 = 13.97 g
Mass of oxygen = 35.89 − 32.37 = 3.52 g

Element	Cu	O
Collected mass (g)	13.97	3.52
Relative atomic mass	64	16
Number of moles	$\frac{13.97}{64} = 0.22$	$\frac{3.52}{16} = 0.22$
Molar ratio (divide by smallest number)	$\frac{0.22}{0.22} = 1$	$\frac{0.22}{0.22} = 1$

Table 9.7 Finding the number of moles of copper and oxygen that reacted

The empirical formula of copper(II) oxide is thus **CuO**.

Molecular Formula

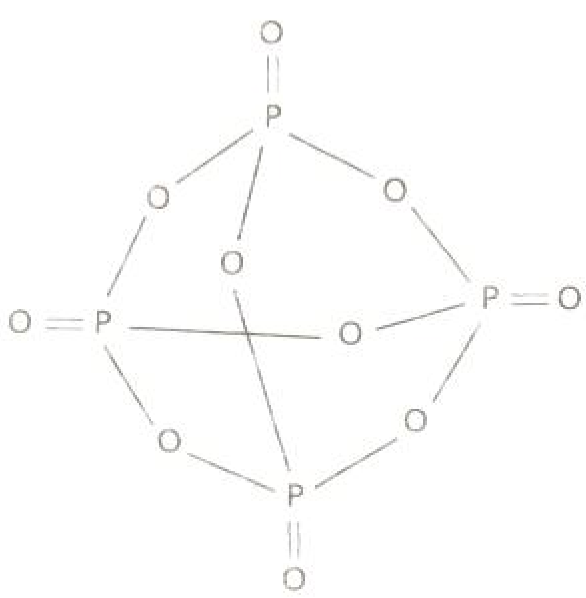


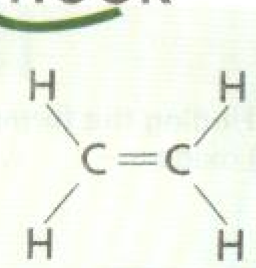
Fig. 9.4 Structure of phosphorus(V) oxide, P_4O_{10}

The empirical formula of phosphorus(V) oxide as determined by experiment is P_2O_5. Its actual formula is P_4O_{10} (Fig. 9.4). We call this the molecular formula. The **molecular formula** is *the formula that shows the exact number of atoms of each element in a molecule.*

What is the relationship between empirical formula and molecular formula?

For many compounds, such as water and ammonia, the empirical formula and molecular formula are the same. However, there are also many compounds (especially organic compounds) whose molecular formulae differ from their empirical formulae. The empirical formulae and molecular formulae of some common compounds are shown in Table 9.8.

Quick Check



What is the empirical formula of this compound?

Substance	Molecular formula	Empirical formula
water	H_2O	H_2O
ammonia	NH_3	NH_3
magnesium oxide	MgO	MgO
hydrogen peroxide	H_2O_2	HO
phosphorus(V) oxide	P_4O_{10}	P_2O_5
ethane	C_2H_6	CH_3

The molecular formula of a compound is a multiple of its empirical formula.

Table 9.8 Empirical formulae and molecular formulae of some common substances

Note that:

1. It is possible for different compounds to have the same empirical formula. For example, ethene (C_2H_4) and propene (C_3H_6) are two compounds with the same empirical formula, CH_2.

2. Where the empirical formula and molecular formula are different, the molecular formula is always a multiple of the empirical formula. For example, the molecular formula and empirical formula of phosphorus(V) oxide are P_4O_{10} and P_2O_5 respectively. The multiple is 2.

We can find the molecular formula of a substance if we know two things: the empirical formula and the relative molecular mass of the substance. They are related as follows:

If empirical formula = A_xB_y, molecular formula = $(A_xB_y)_n$
(where $n = 1, 2, 3$, etc.).
To find n,

$$n = \frac{\text{relative molecular mass}}{M_r \text{ from empirical formula}}$$

Example 1

The empirical formula of ethane is CH_3. Given that the relative molecular mass of ethane is 30, what is its molecular formula?

Solution:

M_r of ethane from empirical formula = 12 + 3 = 15

$$\frac{\text{relative molecular mass}}{M_r \text{ from empirical formula}} = \frac{30}{15} = 2$$

Hence, the molecular formula of ethane = $C_{1\times2}H_{3\times2}$ = $\mathbf{C_2H_6}$

Example 2

Compound *X* contains 40.0% carbon, 6.6% hydrogen and 53.3% oxygen. Its relative molecular mass is 180. What is the molecular formula of *X*?

Solution:

The percentage of each element is directly proportional to its mass in grams. Thus, the mass of each element in 100 g of compound is its percentage in the compound.

Element	C	H	O
Percentage in compound (%)	40.0	6.6	53.3
Relative atomic mass	12	1	16
Number of moles	$\frac{40.0}{12} = 3.3$	$\frac{6.6}{1} = 6.6$	$\frac{53.3}{16} = 3.3$
Molar ratio	$\frac{3.3}{3.3} = 1$	$\frac{6.6}{3.3} = 2$	$\frac{3.3}{3.3} = 1$

The empirical formula of X is thus CH_2O.

M_r of CH_2O = 12 + (2 × 1) + 16 = 30

$$\frac{\text{relative molecular mass}}{M_r \text{ from empirical formula}} = \frac{180}{30} = 6$$

Therefore, the molecular formula of X = $(CH_2O)_6$ = $\mathbf{C_6H_{12}O_6}$

Key ideas

1. The empirical formula shows the types of elements present in the simplest ratio in the compound.
2. The molecular formula shows the exact number of atoms of each element in a molecule.
3. $\frac{\text{relative molecular mass}}{M_r \text{ from empirical formula}} = n$ (where n = 1, 2, 3, etc.)

Test Yourself 9.4

Worked Example

Glycerol contains 39.1% carbon, 52.2% oxygen and the remainder is hydrogen. What is the molecular formula of glycerol? (1 mole of glycerol weighs 92.0 g.)

Answer

Element	C	H	O
Percentage in compound (%)	39.1	100 − 39.1 − 52.2 = 8.7	52.2
Relative atomic mass	12	1	16
Number of moles	$\frac{39.1}{12} = 3.3$	$\frac{8.7}{1} = 8.7$	$\frac{52.2}{16} = 3.3$
Molar ratio	$\frac{3.3}{3.3} = 1$	$\frac{8.7}{3.3} = 2.6$	$\frac{3.3}{3.3} = 1$
Simplest ratio	1 × 3 = **3**	2.6 × 3 = **8**	1 × 3 = **3**

Therefore, the empirical formula of glycerol is $C_3H_8O_3$.

The relative molecular mass from empirical formula
= (3 × 12) + (8 × 1) + (3 × 16)
= 92

$$\frac{\text{relative molecular mass}}{M_r \text{ from empirical formula}} = 1$$

Therefore, the molecular formula of glycerol = $(C_3H_8O_3)_1$ = $\mathbf{C_3H_8O_3}$

Questions

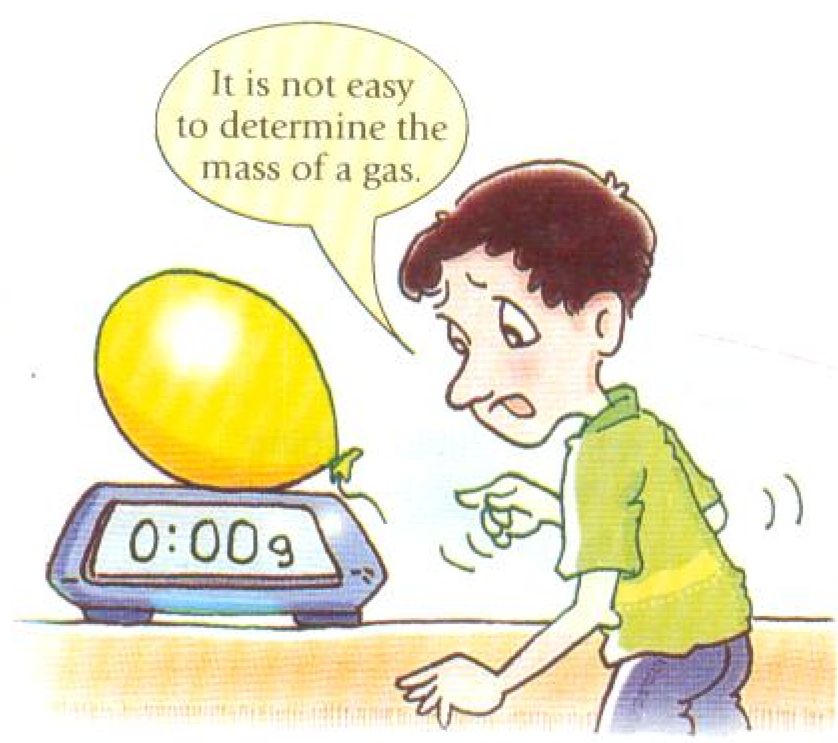
1. Caffeine is a compound found in coffee and tea. The percentage composition of caffeine is 49.5% carbon, 5.1% hydrogen, 16.5% oxygen and 28.9% nitrogen. The relative molecular mass of caffeine is 195. Determine
   a) the empirical formula and
   b) the molecular formula of caffeine.

2. The following results were obtained in an experiment to determine the formula of an oxide of silicon.
   Mass of crucible = 15.20 g
   Mass of crucible + silicon = 15.48 g
   Mass of crucible + oxide of silicon = 15.80 g
   a) Find the empirical formula of the oxide of silicon.
   b) If the M_r of the oxide of silicon is 60, what is its molecular formula?

3. A sample of hydrated copper(II) sulphate weighs 124.8 g. The sample has been determined to contain 31.8 g of copper(II) ions and 48.0 g of sulphate ions.
   a) How many molecules of water of crystallisation are present in the sample?
   b) Deduce the actual formula of hydrated copper(II) sulphate.

9.7 | Molar Gas Volume

You have learnt that the mass of one mole of a substance has a numerical value equal to its relative atomic mass or relative molecular mass. This property is applicable to solids, liquids and gases. However, it is easier to measure the volume of a gas than its mass, since gases weigh very little. Is there a way to relate moles to the volumes of gases?



> **Avogadro's Law** states that equal volumes of all gases, under the same conditions of temperature and pressure, contain the same number of molecules.

In fact, chemists have found by experiment that **one mole of any gas** occupies **24 dm^3** (24 000 cm^3) at room temperature and pressure (r.t.p.). This volume is called the **molar volume** of a gas.

Chem-Aid

r.t.p. or room temperature and pressure are often taken as the conditions of 25 °C and 1 atm.

This means that at r.t.p.
- 1 mol of oxygen occupies 24 dm^3,
- 1 mol of carbon dioxide occupies 24 dm^3,
- 2 mol of oxygen occupy $2 \times 24 = 48$ dm^3,
- 2 mol of carbon dioxide occupy $2 \times 24 = 48$ dm^3.

Does one mole of gas in a hot air balloon occupy 24 dm³?

How can we calculate the number of moles of a gas?

The number of moles of a gas can be measured in two ways.

1. Find the mass of the gas. Then use the following formula to calculate the number of moles of the gas.

$$\text{Number of moles of gas} = \frac{\text{mass of gas in grams}}{M_r \text{ of gas}}$$

2. Find the volume of the gas. Then use this formula:

$$\text{Number of moles of gas} = \frac{\text{volume of gas in cm}^3 \text{ at r.t.p.}}{24\,000 \text{ cm}^3}$$

This formula can be rearranged to give

$$\text{Volume of gas (in cm}^3) = \text{number of moles} \times 24\,000$$
$$\text{Volume of gas (in dm}^3) = \text{number of moles} \times 24$$

Example 1

What is the volume, in dm^3, of 8 g of oxygen gas (O_2) at r.t.p.?

Solution:

Relative molecular mass of oxygen = 2 × 16 = 32

$$\text{Volume of oxygen} = \text{number of moles of oxygen} \times 24$$
$$= \frac{\text{mass of oxygen}}{M_r \text{ of oxygen}} \times 24$$
$$= \frac{8}{32} \times 24$$
$$= \mathbf{6\ dm^3}$$

Example 2

In an experiment, hydrochloric acid was reacted with calcium carbonate at room temperature and pressure (Fig. 9.5). 80 cm^3 of carbon dioxide was produced. Calculate the number of molecules of carbon dioxide given off.

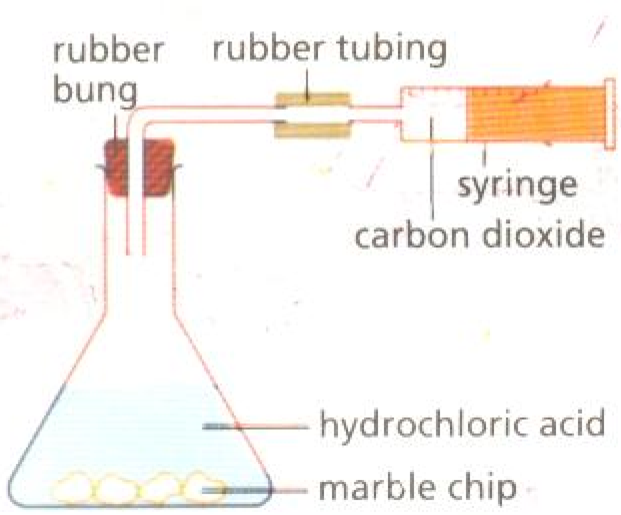


Fig. 9.5 Measuring the volume of carbon dioxide given off

Solution:

Number of moles of carbon dioxide given off

$$= \frac{80}{24\,000}$$
$$= 3.33 \times 10^{-3} \text{ mol}$$

Number of molecules of carbon dioxide given off

= number of moles × Avogadro's constant

$= 3.33 \times 10^{-3} \times 6 \times 10^{23}$

$= \mathbf{2.00 \times 10^{21}}$

Example 3

Calculate the mass of oxygen gas (O_2) in a room that measures 4 m high, 8 m wide and 10 m long. Assume that air contains 20% oxygen. ($1\ m^3 = 10^6\ cm^3$)

Solution:

Volume of air in the room $= 4 \times 8 \times 10$
$= 320\ m^3$
$= 320 \times 10^6\ cm^3$

Volume of oxygen in the room $= (320 \times 10^6) \times 20\%$
$= 64 \times 10^6\ cm^3$

Mass of oxygen $=$ number of moles of oxygen $\times M_r$

$$= \frac{\text{volume of oxygen}}{\text{molar volume}} \times 32$$

$$= \frac{64 \times 10^6\ cm^3}{24\ 000\ cm^3} \times 32$$

$$= \mathbf{8.53 \times 10^4\ g}$$

Do the balloons of the same volume contain the same number of particles?

Look at the balloons on the right. They are of the same volume but do they contain the same number of particles?

Yes, they do! According to Avogadro's Law, each of these balloons contains the same number of gaseous particles since they have the same volume.

Balloons containing identical volumes of gas

Do the balloons of the same mass contain the same number of particles?

If the balloons each contained 0.18 g of a different gas (instead of having the same volume), they would *not* contain the same number of gaseous particles. The following calculation explains why.

Gas	Mass (g)	A_r or M_r	Number of moles $= \frac{\text{mass (g)}}{A_r \text{ or } M_r}$
helium (He)	0.18	4	$\frac{0.18}{4} = 0.045$
hydrogen (H_2)	0.18	2	$\frac{0.18}{2} = 0.090$
methane (CH_4)	0.18	16	$\frac{0.18}{16} = 0.011$

Table 9.9 Calculating the number of moles of different gases with equal mass

Number of helium atoms $= 0.045 \times 6 \times 10^{23} = 2.70 \times 10^{22}$
Number of hydrogen molecules $= 0.090 \times 6 \times 10^{23} = 5.40 \times 10^{22}$
Number of methane molecules $= 0.011 \times 6 \times 10^{23} = 6.60 \times 10^{21}$

Therefore, equal masses of different gases do *not* contain the same number of particles.

Key ideas

1. Avogadro's Law states that equal volumes of all gases under the same conditions of temperature and pressure contain the same number of particles.
2. The volume occupied by one mole of a gas is called its molar volume.
3. At room temperature and pressure, the molar volume of a gas is equal to 24 dm^3 or 24 000 cm^3.

Test Yourself 9.5

Worked Example

Which substance contains 6 × 10^{23} atoms at room temperature and pressure?

(A_r: N = 14; O = 16; Ne = 20; Mg = 24; Al = 27; Cl = 35.5)

A 12 000 cm^3 of neon
B 41 g of aluminium nitride (AlN)
C 12 dm^3 of oxygen
D 40 g of magnesium oxide (MgO)

Thought Process

Oxygen exists as diatomic molecules (O_2). 12 dm^3 of oxygen contain $\frac{12}{24}$ = 0.5 mol of oxygen molecules (O_2). Each molecule of oxygen contains two atoms of oxygen. Therefore, 0.5 mol of oxygen contains 1 mol of oxygen atoms. Hence, there are 6 × 10^{23} oxygen atoms.

Answer

C

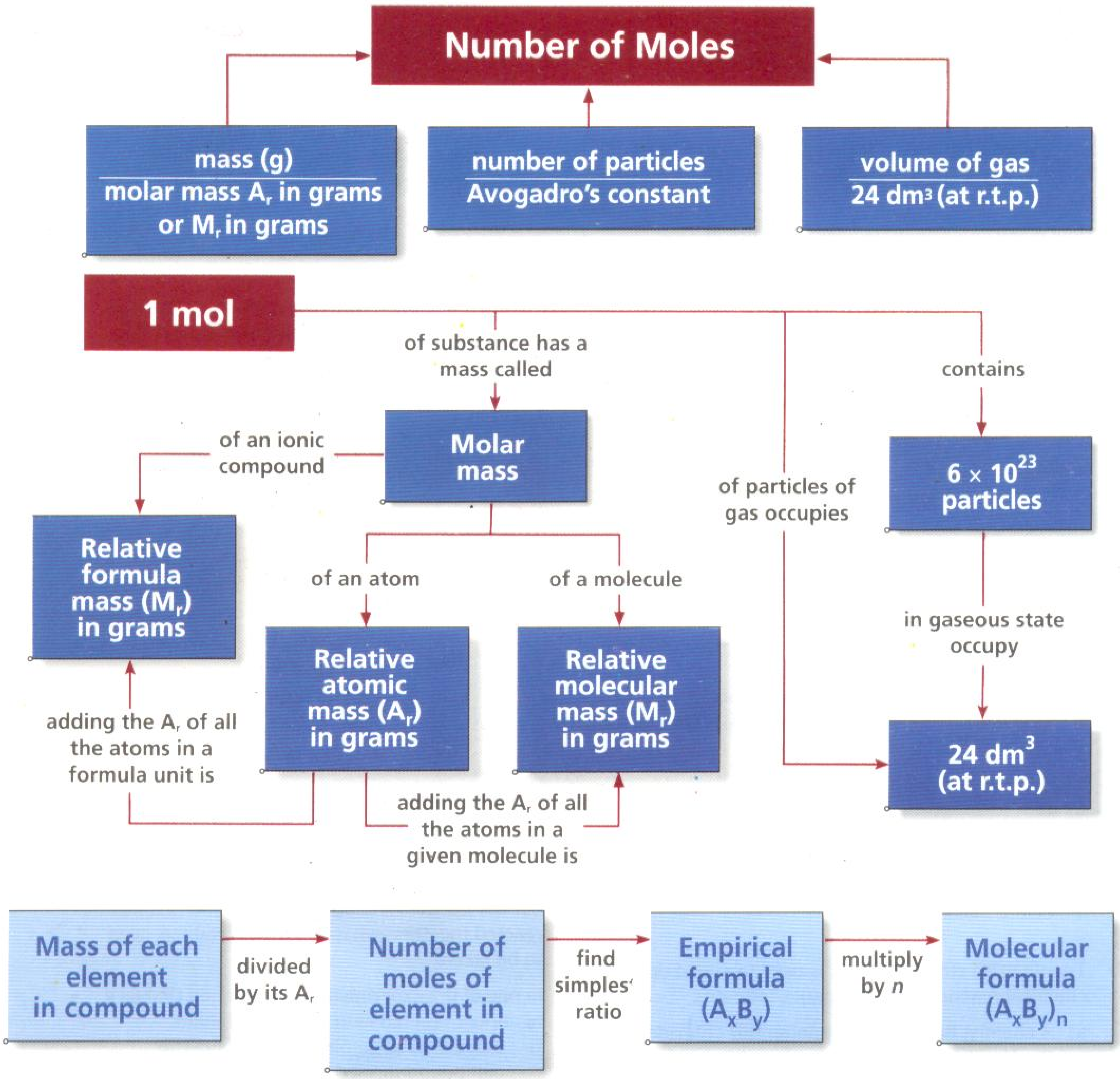
Questions

1. Calculate the number of
   a) sulphur atoms in 64 g of sulphur.
   b) magnesium atoms in 0.01 mol of magnesium.
   c) methane molecules in 112 cm^3 of methane at r.t.p.
   d) atoms in a drop of water weighing 0.5 g.

2. How many moles are present in the following volumes of gases (at r.t.p.)?
   a) 1.2 dm^3 of sulphur dioxide, SO_2
   b) 0.24 dm^3 of methane, CH_4
   c) 120 cm^3 of carbon dioxide, CO_2

3. 0.52 g of a metal (A_r = 65) reacts with 192 cm^3 of chlorine under room conditions to form a compound. What is the formula of the compound?

Concept Map



n = M_r divided by M_r of A_xB_y

Exercise 9

Foundation

1. Which step shows the correct way of calculating the relative formula mass of sodium carbonate crystals, $Na_2CO_3.10H_2O$?

   A $(2 \times 11) + 6 + (3 \times 8) + 10 \times 9$
   B $(2 \times 23) + (3 \times 16) + 10 \times 16$
   C $(2 \times 23) + 12 + (3 \times 16) + (10 \times 2) + (10 \times 16)$
   D $(2 \times 23) + 12 + (3 \times 16) + 10 + 160$

2. How many chlorine molecules are there in 71 g of gaseous chlorine?

   A $0.5 \times 6 \times 10^{23}$ B $2 \times 6 \times 10^{23}$
   C $35.5 \times 6 \times 10^{23}$ D 6×10^{23}

3. What amount, in moles, of calcium oxide is formed when 10.0 g of calcium is burnt in a plentiful supply of oxygen?

   A 0.125 B 0.25
   C 4 D 6

4. Find the empirical formula of a compound which contains 40% sulphur and 60% oxygen by mass.

   A S_2O B SO
   C SO_2 D SO_3

5. An ionic compound of the element *X* and chlorine has the formula *XCl*. The M_r of *XCl* is 74.5. What is the M_r of the oxide of *X*?

   A 55 B 71
   C 87 D 94

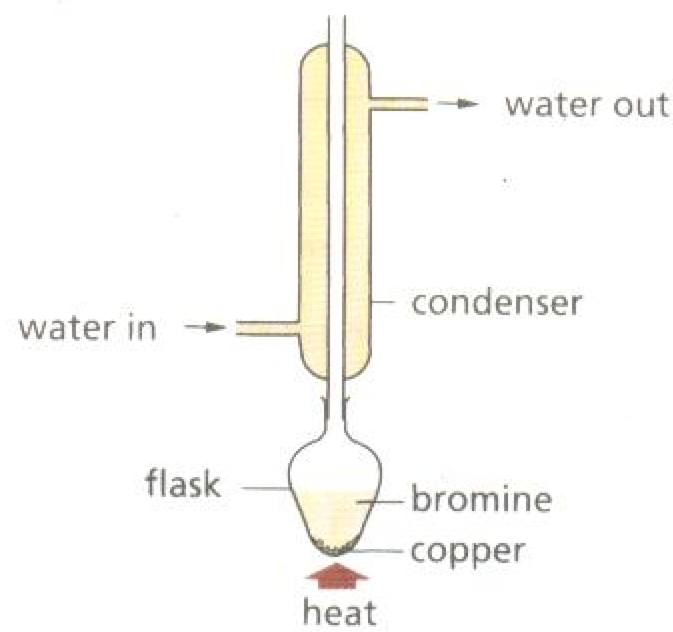
6. A sample of gas weighs 2.0 g and has a volume of 3 dm^3 at room conditions. What is the relative molecular mass of the gas?

7. Titanium(III) chloride, $TiCl_3$, is used to speed up chemical reactions in the plastics industry.
   a) A chemist has 38.6 g of titanium(III) chloride.
      i) How many moles of titanium(III) chloride does the chemist have?
      ii) What is the mass of titanium in this sample?
   b) How many chlorine atoms are there in 0.5 mol of titanium(III) chloride?

8. a) Calculate the percentage by mass of tin in tin(II) chloride ($SnCl_2$).
   b) Tin reacts with chlorine to form another chloride containing 45.6% of tin.
      i) What is the empirical formula of this chloride?
      ii) If the M_r of this chloride of tin is 261, what is the molecular formula of the chloride?

Challenge

A compound of copper and bromine was produced in a fume cupboard using the apparatus shown. Bromine has a melting point of –7 °C and a boiling point of 58 °C.



a) Suggest why
   i) the experiment was performed in a fume cupboard.
   ii) the above apparatus was used.

b) The flask was weighed empty. In order to work out the masses of copper and bromine that had reacted together, what **two** other readings need to be taken?

c) In one of the experiments, it was found that 64 g of copper reacted with 160 g of bromine.
   i) Calculate the empirical formula of the compound formed between copper and bromine.
   ii) What further information is required to find the molecular formula of the compound?