Physical Constants

<table>
<thead>
<tr>
<th>Name</th>
<th>Common Symbol</th>
<th>Value</th>
</tr>
</thead>
<tbody>
<tr>
<td>Avogadro constant</td>
<td>(N_A)</td>
<td>(6.022 \times 10^{23}) (\text{mol}^{-1})</td>
</tr>
<tr>
<td>Boltzmann constant</td>
<td>(k)</td>
<td>(1.3807 \times 10^{-23}) (\text{J} \cdot \text{K}^{-1})</td>
</tr>
<tr>
<td>Elementary charge</td>
<td>(e)</td>
<td>(1.6022 \times 10^{-19}) (\text{C})</td>
</tr>
<tr>
<td>Faraday constant</td>
<td>(F)</td>
<td>(9.6485 \times 10^{4}) (\text{C} \cdot \text{mol}^{-1})</td>
</tr>
<tr>
<td>Gas constant</td>
<td>(R)</td>
<td>(8.3145) (\text{J} \cdot \text{K}^{-1} \cdot \text{mol}^{-1})</td>
</tr>
<tr>
<td>Mass of an electron</td>
<td>(m_e)</td>
<td>(9.1094 \times 10^{-31}) (\text{kg})</td>
</tr>
<tr>
<td>Planck constant</td>
<td>(h)</td>
<td>(6.6261 \times 10^{-34}) (\text{J} \cdot \text{s})</td>
</tr>
<tr>
<td>Speed of light</td>
<td>(c)</td>
<td>(2.9979 \times 10^{8}) (\text{m} \cdot \text{s}^{-1})</td>
</tr>
<tr>
<td>Vacuum permittivity</td>
<td>(\varepsilon_0)</td>
<td>(8.8542 \times 10^{-12}) (\text{C}^2 \cdot \text{J}^{-1} \cdot \text{m}^{-1})</td>
</tr>
</tbody>
</table>

Atomic Mass

The modern system of atomic masses is based on \(^{12}\text{C}\) (carbon 12). In this system, \(^{12}\text{C}\) is assigned a mass of exactly 12 atomic mass units.

The relative atomic mass, \(A_r\) of an element is the average mass of the naturally occurring isotopes of the element relative to the mass of an atom of \(^{12}\text{C}\).

The relative molecular mass, \(M_r\), is the mass of a molecule relative to the mass of an atom of \(^{12}\text{C}\).

e.g. \(A_r\) (H) = 1, \(A_r\) (O) = 16, \(M_r\) (\(\text{CH}_2\)) = 16, \(M_r\) (\(\text{NH}_3\)) = 17

The relative formula mass is used for ionic substances.

Relative masses have NO UNITS.

The Mole

Real samples contain very large numbers of atoms, molecules or ions. To simplify matters, a special unit is defined called the mole. This is defined as the amount of substance that contains the same number of particles as there are atoms in 12 g of \(^{12}\text{C}\).

The number of atoms in 12 g of \(^{12}\text{C}\) is known as the Avogadro number (L).

\[ L = 6.022 \times 10^{23} \]

Thus 1 mole of any substance contains \(6.022 \times 10^{23}\) separate atoms, molecules or formula units of that substance.

e.g. 1 mole of \(\text{O}_2\) contains \(6.022 \times 10^{23}\) oxygen molecules
1 mole of \(\text{Al}\) contains \(6.022 \times 10^{23}\) aluminium atoms

Never forget that the mole is just a convenient unit, like ‘a dozen’.

A dozen eggs is 12 eggs
A mole of eggs is \(6.022 \times 10^{23}\) eggs

The mole is the SI unit for the amount of substance and has the symbol mol.

### Empirical and molecular formulae

The molecular formula of a compound shows the number of atoms of each element that are present in one molecule of the compound.

e.g. molecular formula of methane = \(\text{CH}_4\)
molecular formula of ethanoic acid = \(\text{C}_2\text{H}_4\text{O}_2\)

The empirical formula of a compound is the formula that shows the simplest possible ratio of all the elements in that compound.

e.g. empirical formula of methane = \(\text{CH}_4\)
empirical formula of ethanoic acid = \(\text{CH}_2\text{O}\)

If the masses or percentage masses of each element present in a compound are known, the empirical formula of the compound can be calculated. If the relative molecular mass is also known, the molecular formula can be calculated.

### Calculating an empirical formula

1) Draw a table with the following headings

<table>
<thead>
<tr>
<th></th>
<th>Mass or % mass</th>
<th>Relative atomic mass ((A_r))</th>
<th>Mass or % mass (A_r)</th>
<th>Test ratio</th>
<th>Simplest ratio</th>
</tr>
</thead>
</table>

2) Enter the data given into the first two columns
3) Find the relative atomic masses for the elements and enter them in the third column
4) Perform the calculation in the fourth column (i.e. divide the value in column 2 by that in column 1). 
5) To determine the test ratio, take the smallest number in column 4 and divide each of the column 4 values by this number – write the answers in column 5.
6) Look at the numbers in column 5. Check that they are all whole numbers. If so, copy them into column 6. If any of them are not whole numbers, multiply all of them by the lowest number needed to make them all whole numbers and write these in column 6.
7) Write out the empirical formula by using the numbers in column 6.

**Example**
a hydrocarbon contains 83.72% carbon by mass. Calculate its empirical formula.

Thus the empirical formula is \(\text{C}_3\text{H}_7\)

### Calculating a molecular formula from an empirical formula and a relative mass.

1) Calculate the mass of your empirical formula (e.g. \(\text{C}_3\text{H}_7 = 36+7 = 43\))
2) Divide the relative mass given by the mass calculated in 1). (e.g. if the relative molecular mass is 86, divide 86 by 43 giving 2)
3) The number calculated gives the scaling factor for each atom in the formula (thus for the example the molecular formula would be \(\text{C}_2\text{H}_4\text{O}_2 = \text{C}_6\text{H}_{14}\)).
Reacting masses

When chemical reactions take place, molecules react in the ratios given by the chemical equation. For instance, in the reaction

\[ 2 \text{H}_2(g) + \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(l) \]

2 molecules of hydrogen react with one molecule of oxygen to make 2 molecules of water. As a mole contains exactly the Avogadro number of molecules, it is also true that 2 moles of hydrogen react with one mole of oxygen to make 2 moles of water.

In any situation the number of moles of a substance can be calculated from its mass and Mr, and vice versa. In any situation the number of moles of a substance can be calculated from its mass and Mr, and vice versa. Therefore, the masses of substance that react with each other can be calculated.

Example

7 g of hydrogen is burnt completely in oxygen. What mass of oxygen is used up and what mass of water is produced?

The following steps allow the answer to be easily calculated.

1) Calculate the number of moles of hydrogen using the given mass and the Mr.

\[ \text{Number of moles} = \frac{\text{mass}}{\text{Mr}} = \frac{7 \text{ g}}{2} = 3.5 \text{ moles} \]

So 7 g of water reacts with 56 g of oxygen to make 63 g of water.

2) Using the chemical equation, determine the corresponding numbers of moles of oxygen used and water produced.

From the chemical equation, 2 moles of hydrogen react with 1 mole of oxygen to give 2 moles of water. Therefore 3.5 moles of hydrogen must react with 1.75 moles of oxygen to give 3.5 moles of water.

3) Use the number of moles from 2) along with the relative molecular mass to calculate the mass of oxygen used and the mass of water produced.

**Oxygen:**

- Number of moles = 1.75
- Mass of oxygen used = Number of moles x Mr = 1.75 x 32 = 56 g

**Water:**

- Number of moles = 3.5
- Mass of water produced = Number of moles x Mr = 3.5 x 18 = 63 g

Reactive volumes

When working with gases, it is much easier to measure their volumes than their masses. As long as the pressure and temperature are kept the same, equal volumes of gases contain equal numbers of molecules. This is Avogadro’s law.

In the laboratory, two sets of conditions are commonly used:

- **Standard Temperature and Pressure (STP):** Temperature = 273 K (0°C) Pressure = 1 atm
- **Room Temperature and Pressure (RTP):** Temperature = 298 K (25°C) Pressure = 1 atm

1 mole of ANY gas at STP has a volume of 22.4 dm³.

We can use this to calculate the volume of gases that react together.

Example

1 dm³ of B contains 4.79 x 10⁻³ / 0.025 = 0.1916

The solution is 0.5 M.

### Solutions

When dealing with solutions, a certain amount of solute is dissolved in a certain amount of solvent. Dissolving more solute or using less solvent results in a more concentrated solution. Imagine making coffee. A normal cup of coffee might use 1 spoon of coffee granules in 1 cup of hot water. If two spoons of coffee granules are used, the coffee will be twice as strong. Also, if half a cup of water is used with 1 spoon of coffee granules, this coffee will also be twice as strong.

As usual, it is helpful to work with moles, so in chemistry the concentration of a solution is called the MOLARITY and tells us the number of moles of solute per litre (= dm³) of solution.

**Example**

If 2.8 g of potassium hydroxide (KOH) is dissolved in 100 cm³ of aqueous solution, what is the molarity of the solution?

1) Calculate the number of moles of KOH.

\[ \text{Number of moles} = \frac{\text{mass}}{\text{Mr}} = \frac{2.8 \text{ g}}{56} = 0.05 \text{ moles} \]

2) Write the volume of solution in dm³. Remember that 1 dm³ = 1000 cm³.

\[ \text{Volume of solution} = 100 \text{ cm³} = 0.1 \text{ dm³} \]

3) Use the equation to calculate the molarity.

\[ \text{Molarity} = \frac{\text{number of moles of solute}}{\text{volume of solution (in dm³)}} = \frac{0.05}{0.1} = 0.5 \text{ M} \]

### Titrations

In titrations we use a solution of known concentration (A) to find the unknown concentration of another solution (B) based on a chemical reaction between the two solutions. During the titration, the volumes of A and B that react together are determined. One is fixed, usually by using a pipette and the other is varied, using a burette.

**Example**

25.00 cm³ of a sodium hydroxide solution of unknown concentration is titrated with 0.100 M sulphuric acid solution. The endpoint occurs when 23.95 cm³ of the acid have been added. Determine the concentration of the sodium hydroxide solution.

**Procedure**

1) Summarise the details of the titration

Solution A : Sulphuric acid, H₂SO₄(aq) : 23.95 cm³ of a 0.100 M solution

Solution B : Sodium hydroxide, NaOH(aq) : 25.00 cm³ of a solution of unknown concentration

2) Write down the chemical equation

\[ \text{H}_2\text{SO}_4(aq) + 2\text{NaOH}(aq) \rightarrow \text{Na}_2\text{SO}_4(aq) + 2\text{H}_2\text{O}(l) \]

3) Calculate the number of moles of solution A

\[ \text{number of moles A} = \frac{\text{volume of solution A (in cm³)}}{1000} \times \text{molarity of solution A} = \frac{23.95 \text{ cm³}}{1000} \times 0.100 = 2.395 \times 10^{-3} \text{ moles} \]

4) Use the equation to relate the number of moles of A to the number of moles of B.

Therefore 2.395 x 10⁻³ moles of H₂SO₄(aq) (A) reacts with 2 moles NaOH (B)

5) Use the number of moles B along with its volume to calculate the concentration

\[ \text{number of moles B} = \frac{\text{volume of solution B (in cm³)}}{1000} \times 0.1916 \text{ moles NaOH} \]

From the equation, 1 mole of H₃SO₄(aq) (A) reacts with 2 moles NaOH (B)

Therefore 2.395 x 10⁻³ moles of H₂SO₄(aq) (A) reacts with 4.79 x 10⁻³ moles NaOH (B)

In the question, the lowest precision number is the concentration of sulphuric acid (3 significant figures) so this precision should be used in the answer.

The solution of NaOH is 0.192 M.