Key Concepts in Chemistry
EdExcel 2016 Chemistry topic 1

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The Weald School
Atomic Structure
The structure of the atom

I first came up with the idea of the atom (after the Ancient Greeks). Our understanding of what it looks like has changed over time due to discoveries of subatomic particles:

- **Electron** - negative, mass nearly nothing
- **Proton** - positive, same mass as neutron (\(1\))
- **Neutron** - neutral, same mass as proton (\(1\))

Dalton
## The structure of the atom

<table>
<thead>
<tr>
<th>Particle</th>
<th>Relative Mass</th>
<th>Relative Charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton</td>
<td>1</td>
<td>+1</td>
</tr>
<tr>
<td>Neutron</td>
<td>1</td>
<td>0</td>
</tr>
<tr>
<td>Electron</td>
<td>1/2000 (i.e. 0)</td>
<td>-1</td>
</tr>
</tbody>
</table>

**MASSE NUMBER** = number of protons + number of neutrons

**SYMBOL**

**ATOMIC NUMBER** = number of protons. All atoms of a particular element must have the same number of protons.
Mass and atomic number

How many protons, neutrons and electrons?

H, 1 protons, 1 neutron, 1 electron
B, 11 protons, 5 neutrons, 8 electrons
O, 16 protons, 8 neutrons, 8 electrons
Na, 23 protons, 11 neutrons, 11 electrons
Cl, 35 protons, 17 neutrons, 17 electrons
U, 238 protons, 92 neutrons, 92 electrons
The structure of the atom in more detail

**ELECTRON** - negative, mass nearly nothing

**PROTON** - positive, same mass as neutron ("1")

**NEUTRON** - neutral, same mass as proton ("1")

The nucleus is around 10,000 times smaller than the atom! The diameter of the nucleus is around $10^{-14}$ m.

The nucleus - this contains most of the mass in an atom

Atoms always have the same number of protons and electrons so they are neutral overall. The atom is around 0.1 nm big (i.e. 10-10 m).
An isotope is an atom with a different number of neutrons:

Notice that the mass number is different. How many neutrons does each isotope have?

Each isotope has 8 protons - if it didn’t then it just wouldn’t be oxygen any more.
Strange atomic masses

When you look at a periodic table sometimes the atomic mass is not a whole number. Consider chlorine, for example:

How can an atom have a decimal for its mass?

This is because out of every four naturally occurring chlorine atoms, 3 have a mass of 35 and 1 has a mass of 37 so the average atomic mass is:

\[
(3 \times 35 + 1 \times 37) / 4 = 35.5
\]

Q. Magnesium is often found as \(^{24}\text{Mg}\) or \(^{26}\text{Mg}\). If 79% of magnesium is \(^{24}\text{Mg}\) what is the average atomic mass?

\[
(79 \times 24 + 21 \times 26) / 100 = 24.4
\]
Periodic Table Introduction

How would you arrange these elements into groups?
The periodic table arranges all the elements in groups according to their properties.

**Vertical columns are called GROUPS**

**Horizontal rows are called PERIODS**
1864: John Newlands arranged the known elements in order of atomic mass and found out that every 8th element had similar properties:

```
Li  Be  B  C  N  O  F  Na  Mg  Al
```

1869: Dimitri Mendeleev arranged the known elements in order of mass but he also left in gaps and was able to predict the properties of unknown elements:

```
Li  Be  B  C  N  O  F  Na  Mg  Al
```

When I came up with my Periodic Table I thought that I had arranged elements in order of atomic mass. However, I was wrong because I did not take into account the masses of different isotopes.
### Fact 1: Elements in the periodic table are arranged in order of proton number:

- Lithium (Li) has 3 protons.
- Beryllium (Be) has 4 protons.
The Periodic Table

**Fact 2:** Most of the elements are metals:

These elements are metals - they form “positive ions”

These elements are non-metals - they form “negative ions”

<table>
<thead>
<tr>
<th>Li</th>
<th>Be</th>
<th>Na</th>
<th>Mg</th>
<th>Al</th>
<th>Si</th>
<th>P</th>
<th>S</th>
<th>Cl</th>
<th>Ar</th>
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</thead>
<tbody>
<tr>
<td>K</td>
<td>Ca</td>
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</table>

This line divides metals from non-metals.
Electron structure

Consider an atom of Potassium:

Potassium has 19 electrons. These electrons occupy specific energy levels “shells”...

The inner shell has ___ electrons
The next shell has ___ electrons
The next shell has ___ electrons
The next shell has the remaining ___ electron

Electron structure = 2, 8, 8, 1
Electron structure

Draw the electronic structure of the following atoms:

N

Electron structure = 2,5

Mg

Electron structure = 2,8,2

Ca

Electron structure = 2,8,8,2
**The Periodic Table**

**Fact 3:** Elements in the same group have the same number of electrons in the outer shell (this corresponds to their group number).

<table>
<thead>
<tr>
<th>H</th>
<th>He</th>
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<tbody>
<tr>
<td>Li</td>
<td>Be</td>
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<tr>
<td>Na</td>
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<tr>
<td>K</td>
<td>Ca</td>
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<tr>
<td>Fe</td>
<td>Ni</td>
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<td>Ag</td>
<td>Pt</td>
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<tr>
<td>Br</td>
<td>Kr</td>
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<tr>
<td>I</td>
<td>Xe</td>
</tr>
</tbody>
</table>

*E.g. all group 1 metals have ___ electron in their outer shell*

*These elements have ___ electrons in their outer shells*

*These elements have ___ electrons in their outer shell*
Fact 4: As you move down through the periods an extra electron shell is added:

- E.g. Lithium has 3 electrons in the configuration 2,1
- Sodium has 11 electrons in the configuration 2,8,1
- Potassium has 19 electrons in the configuration __,____,____,____
Ionic Bonding
Compounds are formed when two or more elements are chemically combined. Some examples:

- Methane
- Sodium chloride (salt)
- Glucose

How are these compounds formed? Let's consider two ways - "ionic" and "covalent" bonding.
Hi. My name’s Johnny Chlorine. I’m in Group 7, so I have 7 electrons in my outer shell.

I’d quite like to have a full outer shell. To do this I need to GAIN an electron. Who can help me?
Ionic Bonding

Here comes a friend, Sophie Sodium

Hey Johnny. I’m in Group 1 so I have one electron in my outer shell. I don’t like only having one electron there so I’m quite happy to get rid of it. Do you want it?

Okay

Now we’ve both got full outer shells and we’ve both gained a charge which attracts us together. We’ve formed an IONIC bond.
An ion is formed when an atom gains or loses electrons and becomes charged:

If we “take away” the electron we’re left with just a positive charge:

This is called an ion (in this case, a positive hydrogen ion, also called a cation).
Ionic bonding

This is where a metal bonds with a non-metal (usually). Instead of sharing the electrons one of the atoms “______” one or more electrons to the other. For example, consider sodium and chlorine:

Sodium has 1 electron on its outer shell and chlorine has 7, so if sodium gives its electron to chlorine they both have a ___ outer shell and are ______.

A _______ charged sodium ion (cation)

A __________ charged chloride ion (_____)

Group 1 ________ will always form ions with a charge of +1 when they react with group 7 elements. The group 7 element will always form a negative ion with charge -1.

Words - full, transfers, positively, negatively, metals, anion, stable
Mass and atomic number

How many protons, neutrons and electrons would you find in these ions?

Hydrogen ion, $H^+$

Sodium ion, $Na^+$

Chloride ion, $Cl^-$
The Periodic Table

Looking at their position in the Periodic Table and understanding their electron structure, we can predict the charge of different ions.

For example, group 1 elements all want to lose one electron so they will all form cations with a charge of +1.

What type of ion (and its charge) will elements from groups 2, 6 and 7 form?
Naming compounds

**Rule 1** - When two elements join and one is a halogen, oxygen or sulphur the name ends with _____ide

e.g. Magnesium + oxygen → magnesium oxide

1) Sodium + chlorine
2) Magnesium + fluorine
3) Lithium + iodine
4) Chlorine + copper
5) Oxygen + iron
6) KBr
7) LiCl
8) CaO
9) MgS
10) KF
Naming compounds

Rule 2 – When three or more elements combine and one of them is oxygen the ending is ______ate

e.g. Copper + sulphur + oxygen → Copper sulphate

1) Calcium + carbon + oxygen
2) Potassium + carbon + oxygen
3) Calcium + sulphur + oxygen
4) Magnesium + chlorine + oxygen
5) Calcium + oxygen + nitrogen
6) AgNO₃
7) H₂SO₄
8) K₂CO₃
Some examples of ionic bonds

Magnesium chloride:

$$\text{Mg} + \text{Cl} \rightarrow \text{Mg}^{2+} + 2\text{Cl}^-$$

Calcium oxide:

$$\text{Ca} + \text{O} \rightarrow \text{Ca}^{2+} + 2\text{O}^2-$$
Balancing ions

<table>
<thead>
<tr>
<th>Some common ions:</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium - Na⁺</td>
</tr>
<tr>
<td>Potassium - K⁺</td>
</tr>
<tr>
<td>Magnesium - Mg²⁺</td>
</tr>
<tr>
<td>Ammonium - NH₄⁺</td>
</tr>
<tr>
<td>Chloride - Cl⁻</td>
</tr>
<tr>
<td>Bromide - Br⁻</td>
</tr>
<tr>
<td>Oxide - O²⁻</td>
</tr>
<tr>
<td>Sulphate - SO₄²⁻</td>
</tr>
</tbody>
</table>

**Determine the formula of these compounds:**

<table>
<thead>
<tr>
<th></th>
<th><strong>Answers:</strong></th>
</tr>
</thead>
<tbody>
<tr>
<td>1) Sodium chloride</td>
<td>1) NaCl</td>
</tr>
<tr>
<td>2) Magnesium oxide</td>
<td>2) MgO</td>
</tr>
<tr>
<td>3) Magnesium chloride</td>
<td>3) MgCl₂</td>
</tr>
<tr>
<td>4) Ammonium chloride</td>
<td>4) NH₄Cl</td>
</tr>
<tr>
<td>5) Sodium sulphate</td>
<td>5) Na₂SO₄</td>
</tr>
<tr>
<td>6) Sodium oxide</td>
<td>6) NaO</td>
</tr>
</tbody>
</table>
Giant Ionic Structures

When many positive and negative ions are joined they form a “giant ionic lattice” where each ion is held to the other by strong electrostatic forces of attraction (ionic bonds).

Notice that there is one chlorine ion for every sodium ion. Therefore the formula for this compound is NaCl.
Covalent Bonding
Hi. My name’s Johnny Chlorine. I’m in Group 7, so I have 7 electrons in my outer shell.

I’d quite like to have a full outer shell. To do this I need to GAIN an electron. Who can help me?
Covalent Bonding

Here comes another one of my friends, Harry Hydrogen. Hey Johnny. I’ve only got one electron but it’s really close to my nucleus so I don’t want to lose it. Fancy sharing?

Now we’re both really stable. We’ve formed a covalent bond.
Covalent bonding

Consider an atom of hydrogen:

Notice that hydrogen has just __ electron in its outer shell. A full (inner) shell would have __ electrons, so two hydrogen atoms get together and “_____” their electrons:

Now they both have a _____ outer shell and are more _____. The formula for this molecule is $\text{H}_2$.

When two or more atoms bond and form a molecule by sharing electrons we call it ____________ BONDING.

Words - covalent, 1, 2, share, full, stable
Types of Substance
Dot and Cross Diagrams for Covalent Molecules

Water, $H_2O$:
Dot and Cross Diagrams for Covalent Molecules

Oxygen, $O_2$: 

[Diagram of Oxygen molecule with two oxygen atoms, each represented by a dot and cross symbolizing the covalent bond between them.]
Dot and cross diagrams

**Water, H₂O:**

Step 1: Draw the atoms with their outer shell:

- H
- H
- O

Step 2: Put the atoms together and check they all have a full outer shell:

- H
- O
- H

**Oxygen, O₂:**

- O
- O

Step 1: Draw the atoms with their outer shell:

Step 2: Put the atoms together and check they all have a full outer shell:
Dot and cross diagrams

Nitrogen, $N_2$:

Ammonia $NH_3$:

Methane $CH_4$:

Carbon dioxide, $CO_2$:
Elements and compounds can form many different structures, including:

1) Ionic, like sodium chloride:
   ![Ionic structure diagram]

2) Giant covalent structures, like graphite:
   ![Giant covalent structure diagram]

3) Metallic, like iron:
   ![Metallic structure diagram]

4) Simple covalent molecules, like methane:
   ![Simple covalent structure diagram]
When many positive and negative ions are joined they form a “giant ionic lattice” where each ion is held to the other by strong electrostatic forces of attraction (ionic bonds).

If these ions are strongly held together what affect would this have on the substance’s:

1) Melting point?
2) Boiling point?
3) State (solid, liquid or gas) at room temperature?
When an ionic structure like sodium chloride is dissolved it enables the water to conduct electricity as charge is carried by the ions:
Metals are defined as elements that readily lose electrons to form positive ions. The electrons in the highest shells are delocalised and surround positive ions. These delocalised electrons can be used to conduct electricity. There are a number of ways of drawing this:
Simple Covalent Molecules

Recall our model of a simple covalent compound like hydrogen, H₂:

Hydrogen has a very low melting point and a very low boiling point. Why?

1) The intermolecular forces are very weak so each one of these H₂ molecules doesn’t really care about the others - it’s very easy to pull them apart.

2) When a substance is heated it is the intermolecular forces that are overcome, NOT the covalent bond in each molecule, which is much stronger!

Also, the molecules do not carry a charge so covalent compounds usually do not conduct electricity.
Giant Covalent structures ("lattices")

Notice that giant covalent structures have very different properties to individual covalent molecules:

1. Diamond - a giant covalent structure with a very ____ melting point due to ______ bonds between carbon atoms

2. Graphite - carbon atoms arranged in a layered structure, with free ______ in between each layer enabling carbon to conduct __________ (like metals)

3. Silicon dioxide (sand) - a giant covalent structure of silicon and oxygen atoms with strong _____ causing a high ______ point and it's a good insulator as it has no free electrons

**Words** - melting, high, electrons, bonds, strong, electricity
<table>
<thead>
<tr>
<th>Element/compound</th>
<th>Property</th>
<th>Uses</th>
<th>Why?</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbon - diamond</td>
<td>Very hard</td>
<td>Drill tips</td>
<td>Extremely strong covalent structure</td>
</tr>
<tr>
<td>Carbon - graphite</td>
<td>Soft</td>
<td>Lubricants</td>
<td>Layers “slip” off each other</td>
</tr>
<tr>
<td>Carbon - graphite</td>
<td>Conducts electricity</td>
<td>Making electrodes</td>
<td>Free electrons to carry charge</td>
</tr>
</tbody>
</table>
Graphene is a single layer of carbon atoms and is only one atom thick.

Q. What are the possible uses for graphene?
Carbon can also be used to make structures called “fullerenes” (carbon atoms forming an empty shape). Fullerenes are compounds used for applications such as drug delivery, lubricants, catalysts and nanotubes and they have structures based on carbon atoms forming hexagonal rings:

A “carbon nanotube” – high tensile strength, high electrical conductivity and high thermal conductivity.

“Buckminster fullerene” – the first fullerene to be discovered in 1960.
Polymers are compounds consisting of large molecules containing chains of carbon atoms. For example, take ethene:

**Step 1:** Break the double bond

**Step 2:** Add the molecules together:

This molecule is called POLYETHENE — a common plastic.
Polymers are compounds consisting of large molecules containing chains of carbon atoms. For example:

- **Ethane**
- **Butane**
- **Ethene**
- **Butene**

This double bond means that alkenes have the potential to join with other molecules - this makes them REACTIVE. Alkenes turn bromine water colourless.
Metals and Alloys

Metals are also easy to bend and malleable. This is because the layers slide over each other:

A pure metal:

An alloy is a mixture of metals that causes the metal to behave differently:
Calculations involving Masses
Other ways of drawing covalent bonds

Consider ammonia (NH$_3$):

What are the limitations of each way of drawing these molecules?
Metals vs Non-metals

Recall that most of the elements are metals:

These elements are metals - they:
1) Have high melting points
2) Can conduct electricity
3) Are dense

This line divides metals from non-metals

These elements are non-metals - they:
1) Have low melting and boiling points
2) Don’t conduct electricity
3) Are not very dense
Relative formula mass, $M_r$

The relative formula mass of a compound is the relative atomic masses of all the elements in the compound added together.

E.g. water $H_2O$:

- Relative atomic mass of $O = 16$
- Relative atomic mass of $H = 1$

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

Work out $M_r$ for the following compounds:

1) $HCl$
   - $H = 1$, $Cl = 35$ so $M_r = 36$

2) $NaOH$
   - $Na = 23$, $O = 16$, $H = 1$ so $M_r = 40$

3) $MgCl_2$
   - $Mg = 24$, $Cl = 35$ so $M_r = 24 + (2 \times 35) = 94$

4) $H_2SO_4$
   - $H = 1$, $S = 32$, $O = 16$ so $M_r = (2 \times 1) + 32 + (4 \times 16) = 98$

5) $K_2CO_3$
   - $K = 39$, $C = 12$, $O = 16$ so $M_r = (2 \times 39) + 12 + (3 \times 16) = 138$
Empirical formulae is simply a way of showing how many atoms are in a molecule (like a chemical formula). For example, \( \text{CaO} \), \( \text{CaCO}_3 \), \( \text{H}_2\text{O} \) and \( \text{KMnO}_4 \) are all empirical formulae. Here’s how to work them out:

**A classic exam question:**

Find the simplest formula of 2.24g of iron reacting with 0.96g of oxygen.

**Step 1:** Divide both masses by the relative atomic mass:

- For iron: \( \frac{2.24}{56} = 0.04 \)
- For oxygen: \( \frac{0.96}{16} = 0.06 \)

**Step 2:** Write this as a ratio and simplify:

\( 0.04:0.06 \) is equivalent to \( 2:3 \)

**Step 3:** Write the formula:

2 iron atoms for 3 oxygen atoms means the formula is \( \text{Fe}_2\text{O}_3 \)
Example questions

1) Find the empirical formula of magnesium oxide which contains 48g of magnesium and 32g of oxygen.

MgO

2) Find the empirical formula of a compound that contains 42g of nitrogen and 9g of hydrogen.

NH₃

3) Find the empirical formula of a compound containing 20g of calcium, 6g of carbon and 24g of oxygen.

CaCO₃
Q. How could you use this experiment to deduce the empirical formula of magnesium oxide?
An example of Conservation of Mass

Here's a classic experiment where magnesium is burned in a crucible:

Mass of magnesium and crucible before burning = 78.25g

2Mg + O₂ → 2MgO

Mass of magnesium and crucible after burning = 78.56g

Mass is always conserved in any reaction, so where did this extra mass come from?
Example Questions

For each of the following reactions, state whether or not the mass of the total system should go up or down and explain your answer:

1) Iron + oxygen $\rightarrow$ iron oxide

2) Copper carbonate $\rightarrow$ copper oxide + carbon dioxide

For higher tier, write a balanced equation for each reaction, given that iron oxide is $\text{Fe}_2\text{O}_3$ and copper carbonate is $\text{CaCO}_3$. 
Calculating the mass of a product

E.g. what mass of magnesium oxide is produced when 60g of magnesium is burned in air?

Step 1: READ the equation:

\[ 2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO} \]

Step 2: WORK OUT the relative formula masses \((M_r)\):

\[
2\text{Mg} = 2 \times 24 = 48 \\
2\text{MgO} = 2 \times (24+16) = 80
\]

Step 3: LEARN and APPLY the following 3 points:

1) 48g of Mg makes 80g of MgO
2) 1g of Mg makes \(80/48 = 1.66\)g of MgO
3) 60g of Mg makes \(1.66 \times 60 = 100\)g of MgO

IGNORE the oxygen in step 2 - the question doesn't ask for it
1) When water is electrolysed it breaks down into hydrogen and oxygen:

\[ 2H_2O \rightarrow 2H_2 + O_2 \]

What mass of hydrogen is produced by the electrolysis of 6g of water?

Work out \( M_r \): 
\[ 2H_2O = 2 \times ((2 \times 1)+16) = 36 \quad 2H_2 = 2 \times 2 = 4 \]

1. 36g of water produces 4g of hydrogen
2. So 1g of water produces \( \frac{4}{36} = 0.11g \) of hydrogen
3. 6g of water will produce \( (\frac{4}{36}) \times 6 = 0.66g \) of hydrogen

2) What mass of calcium oxide is produced when 10g of calcium burns?

\[ 2Ca + O_2 \rightarrow 2CaO \]

\( M_r \): 
\[ 2Ca = 2 \times 40 = 80 \quad 2CaO = 2 \times (40+16) = 112 \]

80g produces 112g so 10g produces \( (112/80) \times 10 = 14g \) of CaO

3) What mass of aluminium is produced from 100g of aluminium oxide?

\[ 2Al_2O_3 \rightarrow 4Al + 3O_2 \]

\( M_r \): 
\[ 2Al_2O_3 = 2 \times ((2 \times 27)+(3 \times 16)) = 204 \quad 4Al = 4 \times 27 = 108 \]

204g produces 108g so 100g produces \( (108/204) \times 100 = 52.9g \) of Al₂O₃
A note about volume...

The two most commonly used units of volume in chemistry are the cm³ and the dm³:

1) Convert 1250cm³ into dm³
2) Convert 1cm³ into dm³
3) Convert 0.056dm³ into cm³
4) Convert 1.28dm³ into cm³
A "Mole" in numbers (higher only)

Definition:

A mole of a substance is the relative formula mass of that substance in grams,

For example, 12g of carbon would be 1 mole of carbon...

...and 44g of carbon dioxide \((CO_2)\) would be 1 mole etc...

Q. How many moles are the following?

1. 23g of sodium
   - 1 mol
2. 48g of magnesium
   - 2 mol
3. 36g of carbon
   - 3 mol
4. 28g of iron
   - 0.5 mol
A “Mole” (higher only)

Definition:

A mole of a substance ALWAYS contains the same number of molecules/ions/particles/atoms:

Avogadro’s Constant: 1 mole = $6.02 \times 10^{23}$ molecules

Q. How many moles are the following?

1. How many molecules are in 2 moles of carbon? $1.2 \times 10^{24}$
2. What about 2 moles of magnesium? $1.2 \times 10^{24}$
3. How many molecules are in 46g of sodium? $1.2 \times 10^{24}$
4. How many molecules are in 23g of iron? $3.0 \times 10^{23}$
Molar Calculations (higher only)

\[ \text{No. of moles} = \frac{\text{Mass (g)}}{\text{Molar mass (g/mol)}} \]

\[ N = \frac{m}{M} \]

Some example questions:

1) Calculate the mass of 4 mol of lithium

2) Calculate the mass of 2 mol of sodium

3) Calculate the number of moles in 36g of carbon

4) Calculate the number of moles in 88g of carbon dioxide

5) Calculate the number of moles in 27g of water

\[ \begin{array}{l}
28g \\
46g \\
3 \text{ mol} \\
2 \text{ mol} \\
1.5 \text{ mol}
\end{array} \]
# Harder Molar Calculations (HT only)

No. of moles = \( \frac{\text{Mass (g)}}{\text{Molar mass (g/mol)}} \)  
\[ N = \frac{m}{M} \]

1 mole = \( 6.02 \times 10^{23} \) molecules

Some example questions:

1) How many moles and how many molecules would be in 46g of sodium?

2 moles, \( 1.2 \times 10^{24} \) molecules

2) How many grams would \( 1.806 \times 10^{24} \) molecules of carbon weigh?

36g

3) How many grams would 1 millions molecules of hydrogen weigh?

\( 1.66 \times 10^{-18} \)g
Concentration

Concentration means “how much of a chemical there is in a fixed volume” and can be measured in $g/dm^3$ or $mol/dm^3$.

A solution of low concentration (‘dilute’) vs. A solution of high concentration (‘strong’).
Questions on Concentration

To calculate the concentration of a substance you could use this formula:

\[
\text{Conc.} = \frac{\text{Mass of substance (g)}}{\text{Volume of solvent (dm}^3)}
\]

Calculate, with units, the concentration of the following:

1) A solution of 10g salt in 1dm\(^3\) of water
2) 2g of hydrochloric acid in 500cm\(^3\) of water
3) 10kg of salt in 200dm\(^3\) of water
4) 0.5g of sodium hydroxide in 100cm\(^3\) of water
Calculating the mass of a product using moles (higher only)

Let’s try this question again but using moles:

Step 1: READ the equation:

\[ 2\text{Mg} + O_2 \rightarrow 2\text{MgO} \]

“2 moles of magnesium + 1 mole of oxygen forms 2 moles of magnesium oxide”

Step 2: WORK OUT the relative formula masses (\(M_r\)) of MgO:

\[ 2\text{MgO} = 2 \times (24+16) = 80 \]

Step 3: Apply these steps:

1) 60g of Mg is equal to 1.25 moles (60/48)
2) Therefore we will make 1.25 moles of magnesium oxide
3) Therefore we make 100g of MgO (1.25 moles)
What happens if you use too much of one compound? Example question:

Consider the reaction you have when you burn methane:

\[ CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O \]

A student burns 32g of methane in 72g of oxygen. Which reactant is used up completely?

32g of methane is 2 moles. 72g of oxygen is 2.25 moles of \( O_2 \). Therefore this reaction is limited by the 2 moles of methane - the “limiting reactant” (i.e. the reactant that is not in excess).

Q. How much \( CO_2 \) would we expect to produce?

2 moles (88g)
Example question:

130g of zinc reacts with 146g of hydrochloric acid (HCl) to form 272g of zinc chloride (ZnCl$_2$) and some hydrogen (H$_2$). Answer the following:

1) How much hydrogen was produced? 4g

2) How many moles of each substance were reacted/produced? 1 of Zn, 2 of HCl, 1 of ZnCl$_2$, 1 of H$_2$

3) Write a balanced chemical equation for this reaction.

\[ \text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2 \]