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A Level Chemistry Summer Independent Learning Activity

Welcome to A Level Chemistry, please complete **ALL** of the following tasks ready for your first day at New College. You can print the booklet, write on the PDF file or answer the questions on paper.

NOTE: An important part of this assignment is to **LEARN** the **highlighted definitions and formulae**.

The activity is split into 6 sections.

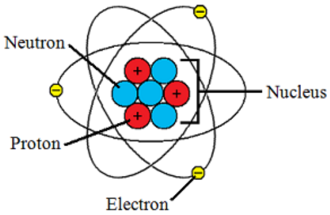
- **TASK 1: Atomic Structure**
- **TASK 2: Chemical Formulae**
- **TASK 3: Writing chemical equations**
- **TASK 4: Mathematical skills**

Please be aware that you will have an assessment on these topics shortly after beginning your A level Chemistry course and the knowledge covered is essential to understanding the subsequent content. Most of the tasks are GCSE revision, but you may need to use secondary sources for small proportion of the work. The following resources may be useful:

- Webpages
  - Chemguide
    http://www.chemguide.co.uk/
  - AQA – Specification at a Glance
    https://www.aqa.org.uk/subjects/science/as-and-a-level/chemistry-7404-7405/specification-at-a-glance
  - RSC
    http://www.rsc.org/learn-chemistry
- YouTube videos
  - MaChemguy
    https://www.youtube.com/user/MaChemGuy/playlists
  - Allery Chemistry
    https://www.youtube.com/channel/UCPtWS4fCi25YHw5SPGdPz0g/playlists?sort=dd&shelf_id=3&view=50
  - Bozeman Science
    https://www.youtube.com/playlist?list=PLllVwaZQkS2op2kDuFifhStNsS49LAxkZ
  - Eliot Rintoul
    https://www.youtube.com/user/MrERintoul
- GCSE notes
- A level and GCSE textbooks

TASK 1: Atomic Structure

This section revises the simple ideas about atomic structure that you will have come across in GCSE. You need to be confident about this before you go on to ideas about the atom which under-pin 'A' level chemistry.



The sub-atomic particles

Protons, neutrons and electrons. – complete the following table

	relative mass	relative charge	position within the atom
proton			
neutron			
electron			

The nucleus

The nucleus is at the centre of the atom and contains the protons and neutrons.

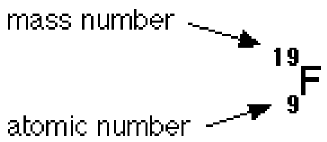
Virtually all the mass of the atom is concentrated in the nucleus, because the electrons weigh so little.

Working out the numbers of protons and neutrons

No of protons = ………….. ………..……….. of the atom

No of protons + no of neutrons = ……….. ………..…..… of the atom

This information can be given simply in the form:



Be careful…sometimes the numbers are shown the other way up!

The atomic number gives the number of protons (9). The atomic number is tied to the position of the element in the Periodic Table.

The mass number counts protons + neutrons (19). If there are 9 protons, there must be 10 neutrons for the total to add up to 19.

Your turn

34 **Se** selenium 79.0

How many protons are in this atom? ……………………………………

How many neutrons are in this atom? ……………………………………

How many electrons are in this atom? ……………………………………

As the atom is neutral in charge, the number of electrons equal the number of protons. If it is an ion, the number of electrons has to be calculate. *Complete the table below.*

Symbol	Atomic no.	Mass no.	Protons	Neutrons	Electrons
$^{79}_{35}Br$					
			12	**13**	**12**
$^{23}_{11}Na^{+}$					
$^{140}_{59}Pr^{2+}$					
			68	**99**	**69**

Isotopes

The number of neutrons in an atom can vary. For example, there are three kinds of carbon atom ^{12}C, ^{13}C and ^{14}C. They all have the same number of protons, but the number of neutrons varies. *Complete the following table.*

	protons	neutrons	mass number
carbon-12	6	6	
carbon-13			13
carbon-14			

These different atoms of carbon are called ***isotopes***. The fact that they have varying numbers of neutrons makes no difference whatsoever to the chemical reactions of the carbon.

Isotopes are ATOMS of the same element, which have the same number of protons (atomic number) but different numbers of neutrons (different mass numbers).

Nickel exists as a mixture of three isotopes, nickel-58, nickel-60 and nickel-62.

Complete the table below to show the atomic structures of the isotopes in metallic nickel.

Isotope	Protons	Neutrons	Electrons
Nickel-58			
Nickel-60			
Nickel-62			

RELATIVE ISOTOPIC MASS

The relative isotopic mass is the mass of an ISOTOPE compared with $^1/_{12}$ of the mass of a carbon-12 atom.

Notice that 'weighted average' is not in this definition. It simply refers to the mass of an atom compared to an atom of carbon -12. It does not take into account the abundance of the isotope.

RELATIVE ATOMIC MASSES

The relative atomic mass of an element is the <u>weighted</u> mean mass of an ATOM of the element compared with $^1/_{12}$ of the mass of a carbon-12 atom.

It has no units because it is a ratio of masses.

Notice that that the **mean mass** of the atoms is used, this is because we take into account the abundance of each isotope of the element that occurs naturally.

The average is a **"weighted mean"** which allows for the fact that there will not be equal amounts of the various isotopes. The example coming up should make that clear:

Suppose you had 100 typical atoms of boron.

19 of these would be ^{10}B and 81 would be ^{11}B.

The total mass of these would be (19 x 10) + (81 x 11) = 1081

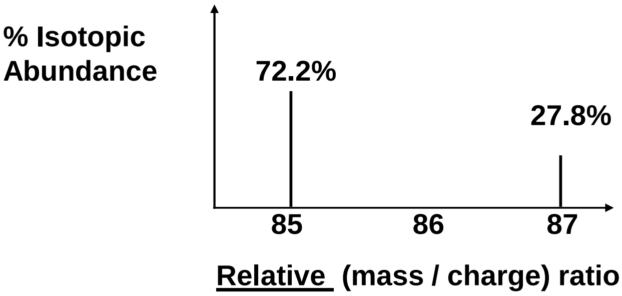
The average mass of these 100 atoms would be 1081/100 = 10.8 (to 3 significant figures).

So 10.8 is the relative atomic mass of boron.

Notice the effect of the "weighted" average. A simple average of 10 and 11 is, of course, 10.5. Our answer of 10.8 allows for the fact that there are a lot more atoms of the heavier isotope of boron.

We can find out the relative isotopic mass and the relative abundances of different isotopes using a MASS SPECTROMETER.

Example 1 : RUBIDIUM



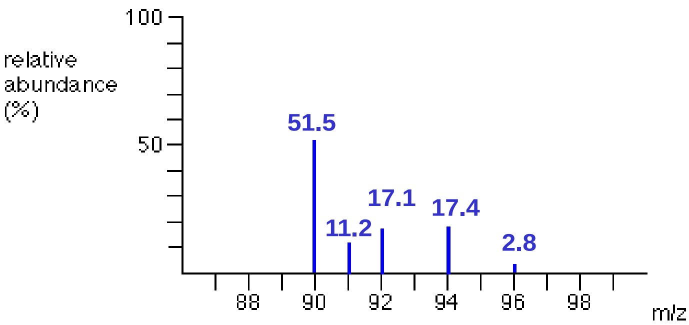
Calculate $A_r(Rb)$ = AVERAGE mass of a Rb atom relative to the mass of a ^{12}C atom

$$= \frac{\text{total relative mass of 100 atoms}}{100}$$

$$= \frac{(72.2 \times 85) + (27.8 \times 87)}{100} = \underline{85.6} \text{ (3sf)}$$

Your turn

Example 2 : ZIRCONIUM



$A_r(Zr)$ =

Calculate the A_r of the following:

a) Chlorine: 75 % ^{35}Cl and 25 % ^{37}Cl

b) Silicon: ^{28}Si, ^{29}Si and ^{30}Si relative abundances 92.18 %, 4.70 % and 3.12 % respectively.

c) Chromium: 4.31 % ^{50}Cr, 83.76 % ^{52}Cr, 9.55 % ^{53}Cr and 2.38 % ^{54}Cr

Exam style question

The antimony in a bullet was analysed by a forensic scientist to help solve a crime. The antimony was found to have the following percentage composition by mass: ^{121}Sb, 57.21%; ^{123}Sb, 42.79%

Calculate a value for the relative atomic mass of the antimony. Give your answer to 4 significant figure.

TASK 2: Chemical Formulae

Ions

Atoms are neutral because they contain the same number of positive protons as negative electrons. For example, the atom Na is neutral because it contains 11 protons (11+ charges) and 11 electrons (11– charges). Ions are particles that contain a different number of protons and electrons. For example, an ion with 11 protons (11+ charges) and 10 electrons (10– charges) has an overall charge of 1^+.

1) Complete the table below to show whether particles are atoms or ions, and for ions their charge.

Number of protons	11+	11+	16+	4+	13+	18+	17+	15+	21+	1+	32+	35+
Number of electrons	11-	10-	18-	2-	10-	18-	18-	18-	18-	0-	32-	36-
Atom or ion?	Atom	Ion	Ion									
Overall charge		1+	2-									

You can also get more complex ions where more than one atom is joined together but has an overall charge. You will have met some of these at GCSE e.g. CO_3^{2-}, SO_4^{2-}.

You will need to use the formulae of ions to write formulae for ionic compounds. You will use the formulae for ionic compounds and molecules to write balanced symbol equations.

You must **LEARN** the common formulae in **bold**. **You will be tested on them in September.**

TABLE OF COMMON FORMULAE

Elements	
Diatomic	
N_2	**Nitrogen**
O_2	**Oxygen**
F_2	**Fluorine**
Cl_2	**Chlorine**
Br_2	**Bromine**
I_2	**Iodine**
H_2	**Hydrogen**
Others	
P_4	Phosphorus
S_8	Sulphur – in practice it is usual to just use S
Most other elements exist as single atoms	

Molecules	
Gases	
NH_3	**Ammonia**
CO_2	**Carbon dioxide**
CO	**Carbon monoxide**
SO_2	**Sulphur dioxide**
CH_4	**Methane**
C_2H_6	Ethane
C_2H_4	Ethene
O_3	Ozone
HCl	Hydrogen chloride
HCN	Hydrogen cyanide
PH_3	Phosphine
Liquids	
H_2O	**Water**
C_2H_5OH	Ethanol
Common acids	
$HCl_{(aq)}$	**Hydrochloric**
HNO_3	**Nitric**
H_2SO_4	**Sulphuric**
H_3PO_4	**Phosphoric**
CH_3COOH	**Ethanoic**

TABLE OF COMMON IONS

In italics are the formulae you can work out using your periodic table.

CATIONS Positively charged ions		ANIONS Negatively charged ions	
Single positive		**Single negative**	
H^+	*Hydrogen ion*	*F^-*	*Fluoride ion*
Li^+	*Lithium ion*	*Cl^-*	*Chloride ion*
Na^+	*Sodium ion*	*Br^-*	*Bromide ion*
K^+	*Potassium ion*	*I^-*	*Iodide ion*
NH_4^+	**Ammonium ion**	**OH^-**	**Hydroxide ion**
Ag^+	**Silver ion**	**NO_3^-**	**Nitrate ion**
Cu^+	*Copper (I) ion*	**HCO_3^-**	**Hydrogencarbonate ion**
		CN^-	**Cyanide ion**
Double positive		**Double negative**	
Mg^{2+}	*Magnesium ion*	*O^{2-}*	*Oxide ion*
Ca^{2+}	*Calcium ion*	*S^{2-}*	*Sulphide ion*
Sr^{2+}	*Strontium ion*	**SO_4^{2-}**	**Sulphate ion**
Ba^{2+}	*Barium ion*	**CO_3^{2-}**	**Carbonate ion**
Zn^{2+}	**Zinc ion**		
Cu^{2+}	*Copper (II) ion*	**$Cr_2O_7^{2-}$**	**Dichromate (VI) ion**
Fe^{2+}	*Iron (II) ion*		
Hg^{2+}	*Mercury (II) ion*		
Pb^{2+}	*Lead (II) ion*		
Triple positive		**Triple negative**	
Al^{3+}	*Aluminium ion*	**PO_4^{3-}**	**Phosphate ion**
Cr^{3+}	*Chromium (III) ion*	*N^{3-}*	*Nitride ion*
Fe^{3+}	*Iron (III) ion*	*P^{3-}*	*Phosphide ion*

Where there are brackets with roman numerals, this is the positive charge on the ion. For example, Iron (II) ion is Fe^{2+} and Iron (III) ion is Fe^{3+}

Use the ion formulae to write formulas for the ionic compounds in the table. Some examples have been done to help you.

	Chloride Cl^-	Oxide	Hydroxide	Sulfate SO_4^{2-}
Sodium				
Magnesium Mg^{2+}	$MgCl_2$			
Iron (III)				$Fe_2(SO_4)_3$
Ammonium NH_4^+		$(NH_4)_2O$		

Name the following ionic compounds:

1) NH_4Cl ____

2) $Fe(NO_3)_3$ ____

3) $TiBr_3$ ____

4) CuS ____

5) $Sn(NO_3)_2$ ____

Write the formulas for the following compounds:

11) Sulphuric acid ____

16) Sodium oxide ____

17) Aluminium hydroxide ____

18) lithium iodide____

19) Calcium sulphate ____

20 Nitric Acid ____________________________________

Name these substances:

Refer to your table of ions where needed

(i) $Mg(NO_3)_2$

(ii) $CuSO_4$

(iii) K_2CO_3

(iv) NH_4OH

(v) Al_2O_3

(vi) $Ca_3(PO_4)_2$

TASK 3: Writing Chemical Equations

Complete these word equations:

magnesium + sulphuric acid → magnesium sulphate + ______________

zinc oxide + magnesium → magnesium oxide + _______________

copper carbonate + sulphuric acid → copper sulphate + water + _____________

strontium hydroxide + nitric acid → ________________ + _______________

Balance these chemical equations:

$$H_2 + Cl_2 \longrightarrow HCl$$

$$H_2 + O_2 \longrightarrow H_2O$$

$$H_2SO_4 + NaOH \longrightarrow Na_2SO_4 + H_2O$$

$$CaCO_3 + HCl \longrightarrow CaCl_2 + CO_2 + H_2O$$

$$HNO_3 + Mg \longrightarrow Mg(NO_3)_2 + H_2$$

$$Al_2O_3 + H_2SO_4 \longrightarrow Al_2(SO_4)_3 + H_2O$$

TASK 4: Mathematical skills

Standard Form

- Standard form is very useful for writing very large or small numbers.
- They are written in the form A x 10^n where A is a number between 1 and 10.
- n represents the number of places the decimal point is moved (for +n values the decimal point has been moved to the left, for -n values the decimal point has been moved to the right).

Number	3435	1029000	0.025	23.2	0.0000278
Standard form	3.435×10^3	1.029×10^6	2.5×10^{-2}	2.32×10^1	2.78×10^{-5}

- To find the value of n:
  - for numbers greater than 1, n = number of places between first number and decimal place
  - for numbers less than 1, n = number of places from the decimal place to the first number (including that number)

Significant figures

Full number	1 sig fig	2 sig fig	3 sig fig	4 sig fig	5 sig fig
9.378652	9	9.4	9.38	9.379	9.3787
4204274	4000000	4200000	4200000	4204000	4204300
0.903521	0.9	0.90	0.904	0.9035	0.90352
0.00239482	0.002	0.0024	0.00239	0.00239	0.002395

Significant figures for calculations involving multiplication / division

- Your final answer should be given to the same number of significant figures as the least number of significant figures in the data used.

  e.g. Calculate the average speed of a car that travels 1557 m in 95 seconds.

  average speed = $\frac{1557}{95}$ = 16 m/s (answer given to 2 sig fig as lowest sig figs in data is 2 sig fig for time)

  e.g. Calculate the average speed of a car that travels 1557 m in 95.0 seconds.

  average speed = $\frac{1557}{95}$ = 16.4 m/s (answer given to 3 sig fig as lowest sig figs in data is 3 sig fig for time)

Significant figures for calculations involving addition/subtraction ONLY

- Here the number of significant figures is irrelevant – it is about the place value of the data. For example

  e.g. Calculate the total energy released when 263 kJ and 1282 kJ of energy are released.

  Energy released = 263 + 1282 = 1545 kJ (answer is to nearest unit as both values are to nearest unit)

  e.g. Calculate the total mass of calcium carbonate when 0.154 g and 0.01234 g are mixed.

  Mass = 0.154 + 0.01234 = 0.166 g (answer is to nearest 0.001 g as least precise number is to nearest 0.001 g)

1) Write the following numbers to the quoted number of significant figures.

a) 345789 4 sig figs ....
b) 297300 3 sig figs ....
c) 0.07896 3 sig figs ....
d) 6.0961 3 sig figs ....
e) 0.001563 3 sig figs ....
f) 0.010398 4 sig figs ....

2) Complete the following sums and give the answers to the appropriate number of significant figures.

a) 6125 x 384 ....
b) 25.00 x 0.010 ....
c) 13.5 + 0.18 ....
d) 7550 ÷ 25 ....
e) 0.000152 x 13.00 ....
f) 0.0125 x 0.025 ....

3) Write the following numbers in non standard form.

a) 1.5×10^{-3} ....
b) 4.6×10^{-4} ....
c) 3.575×10^{5} ....
d) 5.34×10^{2} ....
e) 1.03×10^{6} ....
f) 8.35×10^{-3} ....

4) Write the following numbers in standard form.

a) 0.000167 ....
b) 0.0524 ....
c) 0.000000015 ....
d) 34500 ....
e) 0.62 ....
f) 87000000 ....

5) Complete the following calculations and give the answers to the appropriate number of significant figures.

a) $6.125 \times 10^{-3} \times 3.5$ ....
b) $4.3 \times 10^{-4} \div 7.00$ ....
c) $4.0 \times 10^{8} + 35000$ ....
d) $0.00156 + 2.4 \times 10^{3}$ ....
e) $6.10 \times 10^{-2} - 3.4 \times 10^{-5}$ ....
f) $8.00 \times 10^{-3} \times 0.100 \times 10^{-3}$ ....

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Working Out Relative Molecular Mass (Mr)

You work out the relative molecular mass of a substance by adding up the relative atomic masses (found on the periodic table) of the atoms it consists of. So, for example, to work out the relative molecular mass of water, H_2O, you add the relative atomic masses of two hydrogens and one oxygen.

M_r of H_2O = (2 x 1) + 16 = 18

To work out the relative molecular mass of $CHCl_3$:

M_r of $CHCl_3$ = 12 + 1 + (3 x 35.5) = 119.5

Calculate the relative formula mass of the following substances.

1. F_2 …………………………………………………………………………………………

2. Fe …………………………………………………………………………………………

3. H_2SO_4 ………………………………………………………………………………………

4. Al_2O_3 …………………………………………………………………………………………

5. $Mg(OH)_2$ ………………………………………………………………………………………

6. $Al(NO_3)_3$ ………………………………………………………………………………………

7. $(NH_4)_2SO_4$ ……………………………………………………………………………………

8. $CuCO_3$ …………………………………………………………………………………………

9. $AgNO_3$ …………………………………………………………………………………………

10. NH_4NO_3 ………………………………………………………………………………………

Avogadro’s Number

Since atoms are so small, any sensible laboratory quantity of substance must contain a huge number of atoms: e.g. 1 gram of magnesium contains 2.5 x 10^{22} atoms.

Such numbers are not convenient to work with, so it is necessary to find a unit of "amount" which corresponds better to the sort of quantities of substance normally being measured. The unit chosen for this purpose is the **mole**. The number is chosen so that 1 mole of a substance corresponds to its relative atomic mass (Ar) or its relative molecular mass (Mr) measured in grams.

One mole of carbon-12 has a mass of 12.0g.

One mole of hydrogen atoms has a mass of 1.0g.

One mole of hydrogen (H_2) molecules has a mass of 2.0g.

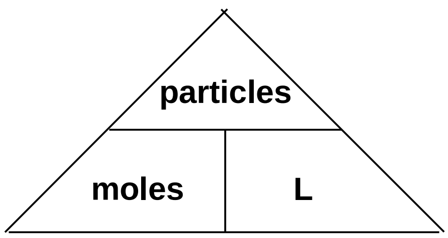
One mole of sodium chloride (NaCl) has a mass of 58.5g.

The number of particles in one mole of a substance is 6.022 x 10^{23}. This is known as **Avogadro's number, L**.

If when we need to know the number of particles of a substance, we usually count the number of moles. It is much easier than counting the number of particles.

The number of particles can be calculated by multiplying the number of moles by Avogadro’s number. The number of moles can be calculated by dividing the number of particles by Avogadro’s number.

Number of particles = Number of Moles x Avogadro’s Constant (L)



Calculations Using Avogadro's Constant

1. One mole of ^{12}C atoms balls have a mass of 12.000 g. Calculate the mass of one ^{12}C atom (give your answer to 3 significant figures).

…………………………………………………………………………………………………

…………………………………………………………………………………………………

2. 5.0 moles of ^{197}Au atoms balls have a mass of 984.8 g. Calculate the mass of one ^{197}Au atom (give your answer to 3 significant figures).

…………………………………………………………………………………………………

…………………………………………………………………………………………………

3. One ^{63}Cu atom has a mass of 1.045×10^{-22} g. Calculate the mass of 3.0 moles of ^{63}Cu atoms (give your answer to 4 significant figures).

…………………………………………………………………………………………………

…………………………………………………………………………………………………

c) One ^{109}Ag atom has a mass of 1.808×10^{-22} g. Calculate the mass of 0.02500 moles of ^{109}Ag atoms (give your answer to 4 significant figures).

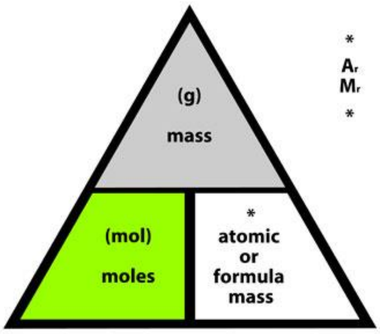
…………………………………………………………………………………………………

…………………………………………………………………………………………………

Working out Moles

The mole is the key concept for chemical calculations.

To work out the mole you use the following equation



$$\text{Number of moles} = \frac{\text{mass in g}}{\text{mass of 1 mole in g}}$$

So, if 1 mole of a substance weighs 40 g and you have 10 g of other words you have 0.25 moles.

Incidentally, the abbreviation for moles which you would normally use in calculations is ***mol***.

Using the Mole Equation

1) Calculate the number of moles of each of the following substances. Give your answers to 3 sig figs.

a) 90.0 g of H_2O ……………………………………………………………………………………

……………………………………………………………………………………………………

b) 20.0 g of C_4H_{10} ………………………………………………………………………………

……………………………………………………………………………………………………

c) 685 g of NH_3 ……………………………………………………………………………………

……………………………………………………………………………………………………

d) 102g of O_2 ………………………………………………………………………………………

……………………………………………………………………………………………………

e) 2.00 kg of Al_2O_3 ……………………………………………………………………………

……………………………………………………………………………………………………

2) Calculate the mass of each of the following substances. Give your answers to 3 sig figs.

a) 4.00 moles of N_2…………………………………………………………………………………

……

b) 0.100 moles of HNO_3………………………………………………………………………………

……

c) 0.0200 moles of K_2O………………………………………………………………………………

……

d) 2.50 moles of PH_3…………………………………………………………………………………

……

e) 0.400 moles of C_2H_5OH……………………………………………………………………………

……

f) 10.0 moles of $Ca(OH)_2$……………………………………………………………………………

……

Working out Empirical Formula

The **empirical formula** of a compound is the formula which shows the simplest whole-number ratio in which the atoms in that compound exist. It can be calculated if the composition by mass of the compound is known.

The **molecular formula** of a substance is the formula which shows the number of each type of atom in the one molecule of that substance.

It applies only to molecular substances and can be deduced if the empirical formula and molar mass of the compound are known. The molecular formula is always a simple whole number multiple of the empirical formula.

Eg a substance contains 85.8% carbon and 14.2% hydrogen, what is its empirical formula? If its relative molecular mass is 56, what is its molecular formula?

Mole ratio = $\frac{85.8}{12}$: $\frac{14.2}{1}$ **$\frac{\text{Mass/percentage}}{\text{Mr}}$**

= $\frac{7.15}{7.15}$: $\frac{14.2}{7.15}$ **Divide your answer by the smallest value to get you lowest whole number ratio**

= 1 : 2 so empirical formula = CH_2 (Mr = 14)

Actual Mr of the molecule = 56

So 56/14 = 4

Molecular formula = C_4H_8

<u>Working Out Empirical Formulas from Masses and Percentage Masses</u>

a) N 82.4%, H 17.6%

……………………………………………………………………………………………………

……………………………………………………………………………………………………

……………………………………………………………………………………………………

b) C 1.24 g H 0.26 g

……………………………………………………………………………………………………

……………………………………………………………………………………………………

……………………………………………………………………………………………………

c) Al 52.9%, O 47.1%

……………………………………………………………………………………………………

……………………………………………………………………………………………………

……………………………………………………………………………………………………

d) Na 0.219 g, H 0.0095 g, C 0.114 g, O 0.457 g

……………………………………………………………………………………………………

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……………………………………………………………………………………………………

e) H 3.1%, P 31.6%, O 65.3%

……………………………………………………………………………………………………

……………………………………………………………………………………………………

……………………………………………………………………………………………………

Working Out Reacting Masses

Step 1 Write ✓ for the substance whose mass is given and ? for the substance whose mass is to be calculated on the balanced equation

Step 2 Find the moles of the ✓ substance (using **moles = mass / Mr**)

Step 3 Use the balanced equation and your answer from step 2 to find the moles of the ? substance

Step 4 Find the mass of the ? substance (using **Mass = moles x Mr**)

Example

What mass of Al_2O_3 will be produced if 10 g of CuO reacts with an excess of Al?

$$3CuO(s) + 2Al(s) \rightarrow Al_2O_3(s) + 3Cu(s)$$

Step 1 ✓ ?

Step 2 moles of CuO = 10/79.5

= 0.126 mol

Step 3 **3:1 ratio with Al_2O_3** so 0.126 / 3 = 0.0419 moles of Al_2O_3

Step 4 mass of Al_2O_3 = 0.0419 x 102 = 4.3 g

Working Out Reacting Masses Questions

1) What mass of oxygen reacts with 12 g of magnesium?

$$2Mg + O_2 \rightarrow 2MgO$$

……………………………………………………………………………………………………

……………………………………………………………………………………………………

……………………………………………………………………………………………………

2) What mass of calcium hydroxide is made from 14 kg of calcium oxide?

$$CaO + H_2O \rightarrow Ca(OH)_2$$

……………………………………………………………………………………………………

……………………………………………………………………………………………………

……………………………………………………………………………………………………

3) What mass of aluminium is needed to react with 640 g of iron oxide?

$$Fe_2O_3 + 2Al \rightarrow 2Fe + Al_2O_3$$

……………………………………………………………………………………………………

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4) What mass of titanium chloride reacts with 460 g of sodium?

$$TiCl_4 + 4Na \rightarrow Ti + 4NaCl$$

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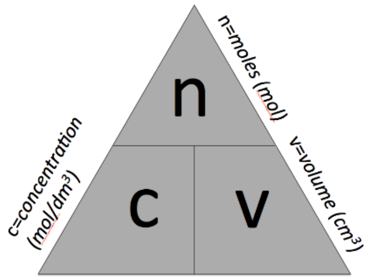
Concentrations of solutions



A **solution** is a mixture of two or more substances.

The major part of a solution is called the **solvent** (usually water) and the minor components are called the **solutes** (solid).
The amount of solute present in a fixed quantity of solvent is called the **concentration** of the solution. It is measured in moles of solute per dm^3 of solution ($moldm^{-3}$).

The following equation is used to work out moles from concentration:

Number of moles = volume x concentration

n = v x c

e.g. if 8 moles are dissolved in 4 dm^3, the concentration is 2.0 $moldm^{-3}$.

How many moles in each of the following?

1) 2 dm^3 of 1.0 $moldm^{-3}$ …………………….

2) 1 dm^3 of 0.25 $moldm^{-3}$…………………….

3) 0.1 dm^3 of 2.5 $moldm^{-3}$…………………….

4) 0.45 dm^3 of 2.0 $moldm^{-3}$…………………….

5) 50 **cm^3** of 0.1 $moldm^{-3}$…………………….

6) 30 **cm^3** of 0.2 $moldm^{-3}$…………………….

What is the concentration of the following solutions?

7) 10 mol dissolved in 2 dm^3 …………………….

8) 2 mol dissolved in 4 dm^3 …………………….

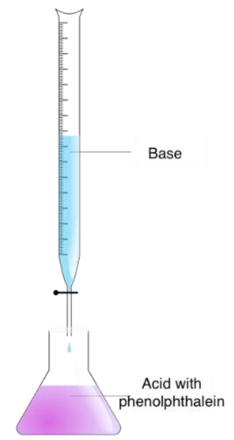
9) 1 mol dissolved in 0.5 dm^3 …………………….

10) 0.145 mol dissolved in 0.1 dm^3 …………………….

11) 0.125 mol dissolved in 25 **cm^3** …………………….

12) 1.75×10^{-3} mol dissolved in 50 **cm^3** …………………….

Working out concentrations in Titrations



Titrations are a very accurate way of measuring the concentration of acids and alkalis.

In a titration, we measure the volume of an acid (or alkali), measured in a burette, needed to exactly neutralise an alkali (or acid) which has been carefully measured into a conical flask with a pipette.

We use an indicator to judge the exact volume required to do this.

1) Place some alkali (or acid) into a conical flask using a pipette.
2) Place the acid (or alkali) into a burette.
3) Add a suitable indicator (e.g. phenol phthalein which works for most titrations)
4) Add the acid (or alkali) from the burette to the conical flask until the colour changes. Do this drop by drop near the end point.
5) Note the final reading.
6) Repeat.

How to do a Titration Calculations

Step 1 Use the volume and concentration of one reactant to calculate the moles.
Step 2 Use the chemical equation to find the moles of the other reactant.
Step 3 Calculate the volume or concentration as required of that reactant.

Note: Volumes are usually given in cm^3 so you need to convert them into dm^3 (divide by 1000)

e.g. 25.0 cm^3 of sulfuric acid reacts with 30.0 cm^3 of 0.150 $moldm^{-3}$ sodium hydroxide. Find the concentration of the acid in $moldm^{-3}$.

$$2\ NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l)$$

Step 1 moles NaOH = conc x vol (dm^3) = 0.150 x 30/1000 = 0.00450 mol

Step 2 2:1 ratio with H_2SO_4
moles H_2SO_4 = moles of NaOH /2 = 0.00450 /2 = 0.00225 mol

Step 3 conc H_2SO_4 = moles / volume = 0.00225 / (25/1000) = 0.0900 $moldm^{-3}$

Titration Questions

1. 25.0 cm^3 of 0.200 $moldm^{-3}$ barium hydroxide solution reacted with 22.8 cm^3 of hydrochloric acid. Calculate the concentration of the hydrochloric acid in $moldm^{-3}$. Give your answer to 3 significant figures.

$$\mathbf{Ba(OH)_2(aq) + 2\ HCl(aq) \rightarrow BaCl_2(aq) + 2\ H_2O(l)}$$

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2. 22.5 cm^3 of sodium hydroxide solution reacted with 25.0 cm^3 of 0.100 $moldm^{-3}$ hydrochloric acid. Calculate the concentration of the sodium hydroxide solution in $moldm^{-3}$. Give your answer to 3 significant figures.

$$\mathbf{NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)}$$

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3. What volume of 0.150 $moldm^{-3}$ rubidium hydroxide reacts with 25.0 cm^3 of 0.240 $moldm^{-3}$ nitric acid? Give your answer to 3 significant figures.

$$\mathbf{RbOH(aq) + HNO_3(aq) \rightarrow RbNO_3(aq) + H_2O(l)}$$

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4. 25.0 cm^3 of 0.200 $moldm^{-3}$ sodium hydroxide solution reacted with 28.7 cm^3 sulfuric acid. Calculate the concentration of the sulfuric acid in $moldm^{-3}$. Give your answer to 3 significant figures.

$$\mathbf{2\ NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l)}$$

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