Calculations Revision materials – Higher

Content will be tested in Chemistry Paper 1 and Paper 2

Checklist

Key points:	☺	☹
Calculations Revision		
How to calculate the mean		
How to calculate % change		
How to work out range		
Work out conversation of units		
Work out standard form		
Relative masses		
How to work out relative formula mass		
Moles		
Reacting masses		
From masses to balanced equations		
Limiting reactants		
How to work out % of an element in a compound		
How to work out concentration		
Bond enthalpies		

Calculating Mean

Key Knowledge

- To calculate the **mean** you **add up** all of the numbers and then **divide** them by how many pieces of data you have.
- If there are any anomalous results, remove them before calculating the mean.

Worked examples

Work out the mean of the data below

12, 13, 14, 11,12, 12, 13

Step 1- Total

12+13+14+11+12+12+13= **87**

Step 2- Divide total by number of pieces of data

87 ÷ 7 = **12.4**

The mean is 12.4

Exam practice 1

1. Calculate the mean reaction time.

person	**1**	**2**	**3**	**4**	**5**
reaction time/seconds	0.258	0.685	0.236	0.246	0.268

Mean =

2. Calculate the mean ADH level in people without diabetes.

people without diabetes insipidus	**ADH level in blood / µg per dm^3**
A	5.2
B	2.8
C	4.9
D	3.5
Mean ADH level:	

Mean ADH levels =

3. Calculate the most appropriate mean volume of oxygen produced at pH 7.

	volume of oxygen produced in cm^3				
pH	**repeat 1**	**repeat 2**	**repeat 3**	**repeat 4**	**mean**
1	1.2	1.6	1.4	1.8	1.5
4	37.7	48.3	38.1	39.9	38.6
7	53.0	51.2	52.8	61.0	
10	29.0	28.5	29.6	28.7	29.3
12	5.2	1.8	1.0	1.4	1.4

Mean at pH 7 =

4. Why do we calculate a mean?

……

……

Calculating percentage change

Key Knowledge

- Equation

$$\frac{\textbf{Final value} - \textbf{starting value}}{\textbf{starting value}} \times 100$$

Worked example

A willow tree initially has a mass of **2.27kg**. After 5 years it has a mass of **76.74kg**

$$\frac{76.74 - 2.27}{2.27} \times 100 = 3281$$

= 3281 % increase

Exam practice 2

1. Calculate the percentage change in mass for chip 5.

chip	concentration of sucrose solution mol per dm^{-3}	starting mass of beetroot chip in grams	end mass of beetroot chip in grams
1	0.0 (water)	2.56	3.89
2	0.2	2.47	2.88
3	0.4	1.99	2.00
4	0.6	2.30	2.12
5	0.8	2.15	1.84
6	1.0	2.22	1.62

2. Suggest why calculating a percentage change is more useful than calculating the change in mass in this investigation

……
……
……

3. Calculate the missing percentage change in mass.

concentration of salt solution / %	mass / g			percentage change / %
	start	after 1 hour	change	
0	10.2	13.1	+2.9	+28.4
10	9.8	11.4	+1.6	+16.3
20	10.3	9.8	−0.5	
30	10.1	8.9	−1.2	−11.9
40	9.7	7.7	−2.0	−20.6

% change =

<u>Range</u>

Key Knowledge

The range is the difference between the highest and lowest values in a set of data.

Worked example

Example 1

Find the range of these numbers: 6, 4, 6, 5, 3.
First put them in order to make it easier to see the lowest and highest.
3, 4, 5, 6, 6

The lowest number is 3 and the highest is 6.

Find the difference. Subtract 3 from 6.

6 - 3 = 3 The range of this set of data is 3.

Exam practice 3

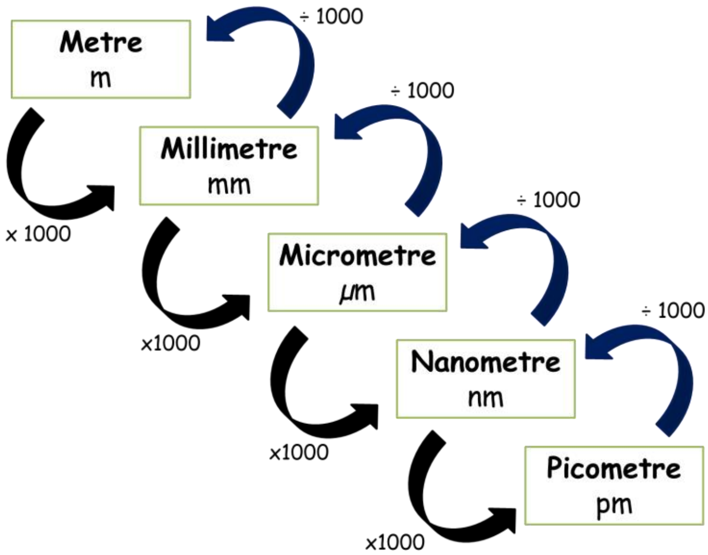
1. Find the range of these numbers: 16, 17, 16,19, 13.

2. Find the range of these numbers: 89, 78, 90, 76, 42.

3. Compare the range of temperatures for Cardiff and London for a week in July.
   Temperatures are given in the table in degrees centigrade.
   Sun Mon Tue Wed Thu Fri Sat

Cardiff 19° 19° 20° 20° 20° 18° 18°
London 20° 22° 22° 21° 20° 21° 19°

<u>**Converting Units**</u>



Worked example

To convert your units do the calculation shown in the diagram
e.g. going from mm to m you divide the number by 1000
6mm = 0.006m

Exam practice 4

1. Convert 34 millimetres into metres.

2. Convert 1034 nanometres into picometres.

3. Convert 6 000 000 seconds into microseconds.

4. Convert 87 nanoseconds into milliseconds.

5. The answer you get to a question is 0.15mm.

   Give your answer in a) m b) µm c) nm

Standard form

Key information

We show figures as numbers between 1 and 10 multiplied by a power of 10
The index number tells us how many place values to move the digit.

Index number

$$A \times 10^n$$

For numbers greater than 0, count how many times you need to move the unit to the right until you form a number between 1 and 10.

Write this number as the power of 10, insert the decimal point and remove the zeros.

1 2 3 4 5 6

$$1150000 = 1.15 \times 10^6$$

You do the reverse for a number smaller than 0 and end up with a negative power.

Exam practice 5

1. Hydrochloric acid with a concentration of 0.001mol/dm^3 is used in a chemical reaction. Give the concentration of the acid in stand form.

   …………………… mol/dm^3

2. In a factory, 134000dm^3 of a chemical are added to a reaction vessel. Write the volume used in standard form

   ……………………dm^3

3. A biologist measures a cell that she is viewing under a microscope.

The width of the cell is 0.00125mm. Write the width in standard form.

……………………mm

Relative masses

Key Knowledge Relative atomic masses

- The mass of a single atom is so tiny that it is not practical to use it in experiments or calculations, that is why we use relative masses.
- For relative atomic masses the carbon-12 ($^{12}_{6}C$) atom is used as the standard atom. The masses of all other atoms are a comparison of their mass to the mass of the carbon-12 atom, e.g. hydrogen has a relative atomic mass of 1, which means most of its atoms have a mass that is $\frac{1}{12}$ of the mass of a $^{12}_{6}C$ atom.
- The symbol for relative atomic mass is A_r
- In the periodic table, the bigger number by each element is its relative atomic mass.
- Relative atomic mass takes into account the relative abundance (proportions) of any isotopes of the element found naturally. That is why some elements have a relative mass with decimals (e.g. chlorine A_r = 35.5).

Worked example:

Chlorine has two principle isotopes, $^{35}_{17}Cl$ and $^{37}_{17}Cl$. Their percentage abundances are 76% and 24%.

To calculate the mean relative mass you need to:

1. Multiply each of the isotopes' masses by their percentage abundance
2. Add the answers from step 1
3. Divide the number from step 2 by 100.

A_r (Cl) = $\frac{(35\times76)+(37\times24)}{100}$ = 35.48 ≈ 35.5

Exam practice 6

1. From the periodic table, find the relative atomic masses of the following elements:
    a. Nitrogen
    b. Magnesium
    c. Argon
    d. Copper
    e. Platinum
    f. Barium
    g. Bismuth
2. Copper has two isotopes, ^{63}Cu and ^{65}Cu. The percentage abundances are 69% and 31%. Calculate the mean relative atomic mass of copper.

Key Knowledge Relative formula mass

- Relative formula mass is the sum of the relative atomic masses of all the atoms shown in a chemical formula of a substance.
- The symbol for relative formula mass is M_r.
- When dealing with molecular substances it can also be referred to as the relative molecular mass.
- You can calculate the percentage by mass of an element in a compound using the atomic masses of the elements and the formula mass of the compound:
    - % by mass(element) = $\frac{A_r(element)\times number\ of\ atoms}{M_r(compound)} \times 100\%$

Worked example, relative formula mass:
Calculate the formula mass of sulfuric acid.
The formula is H_2SO_4. The A_r(H)=1, A_r(S)=32, A_r(O)=16.
M_r = (1x2) + 32 + (16x4) = 2 + 32 + 64 = 98

Worked example, percentage by mass:
Calculate the percentage by mass of hydrogen in water.
A_r(H)=1
A_r(O)=16
$M_r(H_2O)$ = (1x2) + 16 = 18
% by mass(H) = $\frac{1\times 2}{18} \times 100\%$ = 11%

Exam practice 7

1. Calculate the relative formula mass for the following compounds:

    a. Nitric acid
    b. Hydrochloric acid
    c. Water
    d. $CaCO_3$
    e. NaOH

2. Calculate the % by mass of oxygen in the following compounds:

    a. H_2O

    b. NaOH

    c. $CaCO_3$

Exam practice 8

Q1.

(ii) Calculate the relative formula mass of ammonium chloride, NH_4Cl.

(Relative atomic masses: H = 1, N = 14, Cl = 35.5)

....

....

Relative formula mass = ...

(2)

Q2. Iron is an essential part of the human diet. Iron(II) sulfate is sometimes added to white bread flour to provide some of the iron in a person's diet.

(a) The formula of iron(II) sulfate is $FeSO_4$

Calculate the relative formula mass (M_r) of $FeSO_4$

Relative atomic masses: O = 16; S = 32; Fe = 56.

....

....

The relative formula mass (M_r) = ...

(2)

Q3. Toothpastes often contain fluoride ions to help protect teeth from attack by bacteria.

Some toothpastes contain tin(II) fluoride.

This compound has the formula SnF_2 .

(a) Calculate the relative formula mass (M_r) of SnF_2.

Relative atomic masses: F = 19; Sn = 119

....

....

....

Relative formula mass (M_r) = ...

(2)

Key Knowledge The mole

- The mole is a shorthand word to describe an amount of substance.
- Just like a dozen means 12, a pair is 2, and a ream is 500, a mole is 6.02×10^{23} particles.
- 6.02×10^{23} is also known as the Avogadro constant.
- A mole of substance is also its relative atomic or formula mass of that substance in grams. E.g. a mole of carbon has a mass of 12 grams and a mole of water has a mass of 18 grams.
- The relative atomic and formula masses also show the mass of one mole of the substance, therefore the unit for A_r and M_r is grams per mole (g/mol).
- You can calculate the number of moles using the following equations:
  - $n\ (number\ of\ moles) = \frac{mass\ (g)}{A_r}$
  - $n\ (number\ of\ moles) = \frac{mass\ (g)}{M_r}$
- If you have been given the number of moles of a substance you can calculate the mass by rearranging the equation:
  - mass (g) = number of moles x A_r
  - mass (g) = number of moles x M_r

Worked example, calculating moles:
How many moles of sulfuric acid are there in 4.9g of sulfuric acid?
m = 4.9g
$M_r(H_2SO_4)$ = 98 g/mol
n = $\frac{4.9}{98}$ = 0.20 mol

Worked example, masses from moles:
What is the mass of 2 moles of sodium chloride?
M_r(NaCl) = 23 + 35.5 = 58.5 g/mol
n = 2 mol
m = n x M_r = 2 x 58.5 = 117g

Exam practice 9

1. Calculate the number of moles:
   a. of helium in 0.02g of helium.

   b. of sulfur in 9.6g of sulphur atoms.

2. What is the mass of:

a. 50 moles of calcium carbonate ($CaCO_3$)?

b. 0.05 moles of hydrogen (H_2)?

c. 0.6 moles of phosphorous (P_4)?

Key Knowledge Reacting masses

- Balance symbol equations can also be used to calculate the mass of reactants used and products produced.
- Steps:
  - From the periodic table, find the A_r of all the elements involved.
  - Calculate the M_r of all the compounds.
  - Using the M_r of the compounds, find the mass of one mole of each substance.
  - Using the balance symbol equation, multiply the mass of one mole of the substance by the number of moles of that substance involved in the chemical reaction.
- If you have been given the mass of one of the substances and asked to calculate the mass of another the steps change slightly:
  - From the periodic table, find the A_r of the elements in the compounds you are interested in.
  - Calculate the M_r of the compounds.
  - Using the M_r of the compound you have been given the mass for, find how many moles of that substance is involved in the reaction.
  - Using the balanced symbol equation, find the mole ratio between the two substances.
  - Using the ratio and the moles of the given substance calculate the number of moles of the substance you are finding the mass of.
  - Using the number of moles you calculated in step 5 and the M_r of the substance you are finding the mass of, calculate the mass of the substance.

Worked example 1

What masses of reactants and products are involved in the balanced symbol equation:

$$H_2 + Cl_2 \rightarrow 2HCl$$

Solution

To do this, you need to know that the A_r for hydrogen is 1 and the A_r for chlorine is 35.5.

A_r of hydrogen = 1 mass of 1 mole of $H_2 = 2 \times 1 = 2$ g

A_r of chlorine = 35.5 mass of 1 mole of $Cl_2 = 2 \times 35.5 = 71$ g

M_r of $HCl = (1 + 35.5)$ mass of 1 mole of $HCl = 36.5$ g

The balanced equation tells you that one mole of hydrogen reacts with one mole of chlorine to give two moles of hydrogen chloride molecules. So turning this into masses you get:

1 mole of hydrogen molecules, H_2	$= 1 \times 2$	= 2 g
1 mole of chlorine molecules, Cl_2	$= 1 \times 71$	= 71 g
2 moles of hydrogen chloride molecules, 2HCl	$= 2 \times 36.5$	= 73 g

Worked example 2

Sodium hydroxide reacts with chlorine gas to make bleach. This reaction happens when chlorine gas is bubbled through a solution of sodium hydroxide. The balanced symbol equation for the reaction is:

2NaOH	+	Cl_2	→	NaOCl	+	NaCl	+	H_2O
sodium hydroxide		chlorine		bleach		salt		water

If you have a solution containing 100.0 g of sodium hydroxide, what mass of chlorine gas do you need to convert it to bleach?

Solution

A_r values: hydrogen = 1, oxygen = 16, sodium = 23, chlorine = 35.5

Mass of 1 mole of	
NaOH	Cl_2
= 23 + 16 + 1 = 40 g	= 35.5 × 2 = 71 g

The table shows that 1 mole of sodium hydroxide has a mass of 40 g.

So 100.0 g of sodium hydroxide is $\frac{100}{40} = 2.5$ moles.

The balanced symbol equation tells you that for every 2 moles of sodium hydroxide you need 1 mole of chlorine to react with it.

So you need $\frac{2.5}{2} = 1.25$ moles of chlorine.

The table shows that 1 mole of chlorine has a mass of 71 g.

So you will need $1.25 \times 71 = 88.75$ g of chlorine to react with 100.0 g of sodium hydroxide.

The answer 88.75 g is given to 4 significant figures. This is to be consistent with the data supplied in the question, as you started with 100.0 g of sodium hydroxide.

The number of significant figures to which the relative atomic masses are quoted does not need to be taken into account in chemical calculations.

Exam practice 10

Q1. The balanced symbol equation for the reaction is

$$H_2\,(g) \quad + \quad Cl_2\,(g) \rightarrow \quad 2HCl\,(g)$$

Starting with 2 g of hydrogen, what mass of hydrogen chloride would be produced?
(Relative atomic masses: H = 1; Cl = 35.5)

....

....

....

....

Mass of hydrogen chloride = ... g

(3 marks)

Q2. Titanium is a transition metal used as pins and plates to support badly broken bones. Titanium is extracted from an ore that contains the mineral titanium oxide. This oxide is converted into titanium chloride. Titanium chloride is heated with sodium to form titanium metal. This reaction takes place in an atmosphere of a noble gas, such as argon.

$$4Na(s) \; + \; TiCl_4(l) \; \rightarrow \; Ti(s) \; + \; 4NaCl(s)$$

Calculate the mass of titanium that can be extracted from 570 kg of titanium chloride.

Relative atomic masses: Cl 35.5; Ti 48.

....

....

....

....

Mass of titanium = ... kg

(3 marks)

Q3. Ammonium nitrate is an important fertiliser. It is made by reacting nitric acid with the alkali ammonia.

(i) State the type of reaction taking place.

....

(1)

(ii) The equation for this reaction is:

$NH_3 + HNO_3 \rightarrow NH_4NO_3$

Calculate the number of tonnes of ammonium nitrate that can be made from 68 tonnes of ammonia.

(Relative atomic masses: H = 1, N = 14, O = 16)

....

....

....

....

(3)
(4 marks)
(Total 10 marks)

From masses to balanced equations

Key Knowledge From masses to balanced equations

- If you have been given the masses of the reactants and the products in a chemical reaction, you can use them to work out the ratio of each of the substances.
- The simplest whole number ratio gives you the balanced equation.
- Steps:
  - Calculate the mass of any missing reactant or product using the conservation of mass.
  - Calculate the M_r for each substance involved in the reaction.
  - Using the M_r and masses, calculate the number of moles of each substance.
  - Find the simplest whole number ratio of the number of moles.
  - Use the ratio to balance the equation.

Worked example:

$$\_\_\ Al + \_\_\ H_2SO_4 \rightarrow Al_2(SO_4)_3 + \_\_\ H_2$$

Mass = 108 $\cancel{g}$	Mass = 588 $\cancel{g}$	Mass = 684 $\cancel{g}$	Mass = 12 $\cancel{g}$
M_r (Al) = 27 $\cancel{g}$/mol	M_r (H_2SO_4) = 98 $\cancel{g}$/mol	M_r ($Al_2(SO_4)_3$) = 342 $\cancel{g}$/mol	M_r (H_2) = 2 $\cancel{g}$/mol
Mol = $\dfrac{4\ mol}{2\ mol}$	Mol = $\dfrac{6\ mol}{2\ mol}$	Mol = $\dfrac{2\ mol}{2\ mol}$	Mol = $\dfrac{6\ mol}{2\ mol}$
2	3	1	3

Exam practice 11

a

$$\_\_\ NO \rightarrow \_\_\ N_2 + \_\_\ O_2$$

Mass = **120 g**	Mass = **56 g**	Mass = **64 g**
M_r (NO) =	M_r (N_2) =	M_r (O_2) =
Mol =	Mol =	Mol =

b Q2

$$\_\_\ HNO_3 + \_\_\ Ca \rightarrow \_\_\ Ca(NO_3)_2 + \_\_\ H_2$$

Mass = **378 g**	Mass = **120 g**	Mass = **492 g**	Mass = **6 g**
M_r (HNO_3)= **63 g/mol**	M_r (Ca) = **40 g/mol**	M_r ($Ca(NO_3)_2$) =**164 g/mol**	M_r (H_2) = **2 g/mol**
Mol =	Mol =	Mol =	Mol =

c Q3

$$\_\_\ Cl_2 + \_\_\ KI \rightarrow \_\_\ KCl + \_\_\ I_2$$

Mass = **71 g**	Mass = **332 g**	Mass = **149 g**	Mass = **254 g**
M_r (Cl_2)= **71 g/mol**	M_r (KI) = **166 g/mol**	M_r (KCl) = **74.5 g/mol**	M_r (I_2) = **254 g/mol**
Mol =	Mol =	Mol =	Mol =

Q4

__ C_3H_8	+ __ O_2	$\rightarrow$ __ CO_2	+ __ H_2O
Mass = **88 g**	Mass = **320 g**	Mass = **264 g**	Mass = ____ **g**
M_r (Cl_2)= **44 g/mol**	M_r (KI) = **32 g/mol**	M_r (KCl) = **44 g/mol**	M_r (I_2) = **18 g/mol**
d Mol =	Mol =	Mol =	Mol =

Key Knowledge Limiting reactants

- In experiments we usually don't use exact amount of substances. One of the reactants will be in excess.
- The reactant that gets used up first is called the limiting reactant.
- The limiting reactant determines the amount of products that can be formed.
- You can use a balanced symbol equation to work out the limiting reactant if you know the masses of the reactants you start with.
- Steps:
  1. Calculate the M_r or A_r of the reactants.
  2. Using the masses given, calculate the number of moles of each reactant.
  3. Check the balance symbol equation to find the mole ratio.
  4. Compare the number of moles of each reactant to the mole ratio to figure out which reactant is limiting.

Worked example:

Worked example 2

If you have 4.8 g of magnesium ribbon reacting in a solution of dilute hydrochloric acid containing 7.3 g of HCl, which reactant is the limiting reactant?

(A_r values: Mg = 24, H = 1, Cl = 35.5)

Solution

The balanced equation for the reaction is:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

You are only interested in the reactants in this question.

number of moles = $\frac{\text{mass}}{A_r}$ or $\frac{\text{mass}}{M_r}$

You start with 4.8 g of Mg, which is $\frac{4.8}{24}$ moles = 0.2 mol

and 7.3 g of HCl, which is $\frac{7.3}{(1 + 35.5)}$ moles = $\frac{7.3}{36.5}$ = 0.2 mol

From the balanced equation, you see that 1 mole of Mg will react with 2 moles of HCl.

Therefore 0.2 mol of Mg will need 0.4 mol of HCl to react completely.

In this case, we have not got 0.4 mol of HCl – we only have 0.2 mol – so the dilute hydrochloric acid is the limiting reactant (and the magnesium is in excess).

Exam practice 12

1 Iron(III) oxide (Fe_2O_3) is reduced by carbon on heating to give iron metal (Fe) and carbon dioxide (CO_2).
When 480 g of Fe_2O_3 is heated with carbon, 336 g of Fe and 198 g of CO_2 are produced.

a Use the law of conservation of mass to work out the mass of carbon that reacted.

....

(*1 mark*)

b Calculate the simplest whole number ratio of moles of Fe_2O_3, C, Fe, and CO_2.

....

....

....

....

....

(*4 marks*)

c Write a balanced equation for the reaction.

....

(*1 mark*)

2 The reaction between copper oxide and carbon can be used to make copper metal.
The equation for this reaction is:

$$2CuO(s) + C(s) \rightarrow 2Cu(s) + CO_2(g)$$

A mixture of 4.0 g of CuO and 1.2 g of carbon is heated.

a Calculate the number of moles in 4.0 g of CuO.

....

....

(*1 mark*)

b Calculate the number of moles in 1.2 g of C.

....

....

(*1 mark*)

c The balanced equation tells us that for every 1 mole of carbon we need 2 moles of copper oxide. Use your answers to **a** and **b** to work out which reactant is the limiting reactant.

....

....

....

(*2 marks*)

d What mass of Cu would you expect to make?

....

....

....

(*2 marks*)

3 The reaction between zinc carbonate and hydrochloric acid can be used to make zinc chloride.
The equation for this reaction is:

$$ZnCO_3(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l) + CO_2(g)$$

6.25 g of $ZnCO_3$ was added to a solution containing 1.825 g of HCl.

a Which reactant is in excess? Show your reasoning.

....

....

....

....

(*4 marks*)

b What mass of zinc chloride would you expect to make?

....

....

(*2 marks*)

Key knowledge for % of an element in a compound

1. Write down the formula of the compound
2. Use the relative atomic mass (A_r) of the elements to calculate the relative formula mass (M_r).
3. Write the mass of the element you are investigating as a fraction of the total M_r.
4. Find the percentage by multiplying the fraction by 100.

Example

What percentage of carbon dioxide is actually carbon?

Formula of carbon dioxide = CO_2

1. Ar of carbon = 12

   Ar of oxygen = 16

   Therefore Mr = 12 + (16 x 2) = 44

2. $\dfrac{\text{Mass of carbon}}{\text{total mass of compound}} = \dfrac{12}{44}$

4. The percentage of carbon in the compound is:

$$\frac{12}{44} \times 100 = 27.3\ \%$$

Exam practice 13

Q1.

(a) The percentage by mass of oxygen in carbon dioxide (CO_2) is calculated by the equation:

$$\text{percentage by mass} = \frac{\text{number of atoms of O} \times \text{Relative atomic mass of oxygen (O)}}{\text{relative molecular mass of carbon dioxide } (CO_2)} \times 100$$

Relative atomic masses (A_r): C = 12 O = 16

Calculate the percentage by mass of oxygen in carbon dioxide (CO_2).

____

____

____

____

Percentage by mass of oxygen = ____________________ %

(3)

Q2.

Some students investigated magnesium oxide.

(a) Magnesium oxide has the formula MgO.

Calculate the percentage by mass of magnesium in magnesium oxide.

____

____

Percentage by mass of magnesium in magnesium oxide = ________%

(2)

Q3.

(a) Molecular formula of copper oxide is CuO.

Calculate the percentage of copper in copper oxide.

____

____

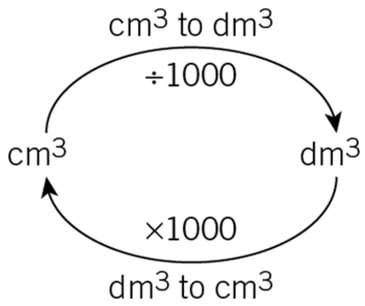
____

Percentage of copper = ____________________ %

(2)

Expressing concentrations

Key Knowledge Volume

- Volume is the amount of space a substance or object occupies.
- There are many units that can be used to measure volume:
    - Millilitres – ml
    - Litres – l
    - Cubic metres – m^3
    - Cubic centimetres – cm^3
    - **Cubic decimetres – dm^3 ← THIS IS THE UNIT CHEMISTS USE MOST**
- 1 litre is the same as 1 dm^3
- 1 ml is the same as 1 cm^3
- Just as there are 1000 millilitres in a litre, there are **1000 cm^3 in 1 dm^3**
- 

Worked examples:

Q1. A solution has a volume of 500cm^3, what is its volume in dm^3?

- 1 dm^3 = 1000 cm^3
- So to convert cm^3 into dm^3 you just have to divide by 1000!

$$\frac{500\ cm^3}{1000} = 0.5\ \text{dm}^3$$

Q2. A solution has a volume of 0.432 dm^3, what is its volume in cm^3?

- 1 dm^3 = 1000 cm^3
- So to convert dm^3 into cm^3 you just have to multiply by 1000!

$$0.432\ \text{dm}^3 \times 1000 = 432\ \text{cm}^3$$

Exam practice 14

1. What is 0.025 dm^3 in cm^3?

2. What is 270 cm^3 in dm^3?

3. How many cm^3 are in 0.052 dm^3?

4. A solution has a total volume of 986 cm^3, what is this in cubic decimetres?

5. 25 cm^3 is taken from a solution of total volume 0.45 dm^3 what volume of the solution remains?

Key Knowledge Concentration

- Solute – the substance that is dissolved in a liquid.
- Solvent – a liquid in which a substance is dissolved.
- Solution – a mixture of the dissolved solute and solvent.
- Concentration – the amount of substance in a certain amount of solution.
- Calculating concentration:
  - $concentration, c\ (g/dm^3) = \frac{amount\ of\ solute, m\ (g)}{volume\ of\ solution, V\ (dm^3)}$
- If the concentration is high, we call the solution concentrated.
- If the concentration if low, we call the solution dilute.
- Increasing the volume (adding more solvent), decreases the concentration.
- Decreasing the volume (evaporating some of the solvent), increases the concentration.
- By rearranging the concentration equation, you can calculate how much solute is in the solution, if you know the concentration and the volume of the solution.

Worked examples:
Q1. If 4g of sodium is dissolved in 2 dm^3 what is the concentration?

$$c\ (g/dm^3) = \frac{m\ (g)}{V\ (dm^3)}$$

c = 4/ 2= 2 g/dm^3

Q2. If 5.5g of sodium hydroxide is dissolved in 3 dm^3 what is the concentration?

c = 5.5/3= 1.83 g/dm^3

Q3. A solution of potassium hydroxide has a concentration of 10 g/dm^3, what mass of potassium hydroxide is dissolved in 0.5 dm^3 of it?

$m\ (g) = c\ (g/dm^3) \times V\ (dm^3)$ m = 10 x 0.5 = 5 g

Exam practice 15

1. If 5 g of sodium is dissolved in 1 dm^3 what is the concentration?

2. If 8 g of sodium hydroxide is dissolved in 2 dm^3 what is the concentration?

3. What is the concentration of a solution when 50 g of hydrogen chloride is dissolved in 5 dm^3 of water?

4. How concentrated is a 500 cm^3 solution that contains 6 g of potassium hydroxide?

5. What mass of lithium is dissolved in 0.4 dm^3 of a 15 g/dm^3 solution?

6. How many grams of hydrogen chloride are dissolved in 4 dm^3 of a 1.4 g/dm^3 solution?

7. A flask contains 400 cm^3 of a 5 g/dm^3 solution of potassium hydroxide. If all the water was evaporated, what mass of potassium hydroxide would remain?

8. A student took 25 cm^3 of 0.1 g/dm^3 sodium thiosulfate solution. What mass of sodium thiosulfate does it contain

Exam practice 16

1. Calculate the concentrations of each of the following solutions in units of g/dm3:
   a) 10.0 g of sodium chloride dissolved in 2.00 dm^3 of water

   ………

   ………

   (1 mark)

   b) 2.5 g of glucose dissolved in 0.5 dm^3 of water

   ………

   ………

   (1 mark)

   c) 3.8 g of copper sulfate dissolved in 250 cm^3 of water

   ………

   ………

   (2 marks)

   d) 25.6 g of potassium chloride dissolved in 1500 cm3 of water.

   ………

....

(2 marks)

2. Calculate the mass of solute dissolved in each of the following solutions in g:
   a) 2 dm^3 copper sulphate solution of concentration 3 g/dm^3.

   ....

   ....

   (1 mark)

   b) 5 dm^3 sodium carbonate solution of concentration 2.5 g/dm^3.

   ....

   ....

   (1 marks)

   c) 250 cm3 copper sulfate solution of concentration 1.2 g/dm^3.

   ....

   ....

   (2 marks)

Bond Enthalpies

Key Knowledge

- All chemicals have chemical energy stored within their chemical bonds. We can find out about the strength of a chemical bond from its bond enthalpy.

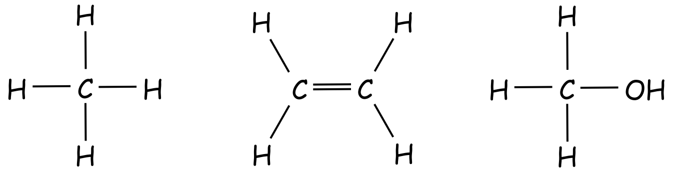
  Bond enthalpies tell you how much energy is required to break different bonds.

  The stronger the bond, the more positive the bond enthalpy.

Breaking bonds is an endothermic process – energy is required.
Making bonds is an exothermic process – energy is released.

The same amount of energy is released when making a new bond as is required for breaking the same bond. For example:
Breaking the bond between two hydrogen atoms in a molecule
$H\text{-}H(g) \rightarrow 2H(g)$ ΔH = +436 kJ mol^{-1} (ENDOTHERMIC)
Making this bond from 2 hydrogen atoms The amount of
energy is the same!
$2H(g) \rightarrow H\text{-}H(g)$ ΔH = -436 kJ mol^{-1} (EXOTHERMIC)

The bond enthalpy for H-H is always the same as this bond can only exist in a H_2 molecule however, some bonds can occur in a number of different molecules. For example, the C-H bond exists in a huge number of organic molecules.



All of these molecules contain C-H bonds, but they are in different environments. The strength of the C-H bond varies across the different environments so average bond enthalpies are used, taking into account the enthalpy of the C-H bond in a wide range of different chemicals.

Exothermic and endothermic reactions

For a chemical reaction to happen bonds between the reactants have to be broken and new bonds are formed. Whether a reaction is exothermic or endothermic depends on the strengths of the bonds being broken and the new bonds being made.

In an endothermic reaction the bonds being broken are stronger than those being made – it requires more energy for the bonds to be broken than making new bonds releases.

In an exothermic reaction the bonds being made are stronger than those being broken – making new bonds releases more energy than is required to break the bonds to begin with.

Energy required to break bonds = Σ(bond enthalpies of bonds broken)

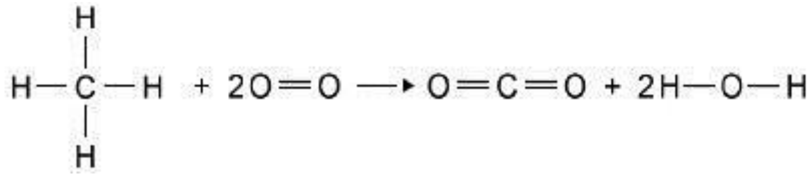
Energy required to break bonds = Σ(bond enthalpies of bonds made)

$\Delta H = \Sigma$(bond enthalpies of bonds broken) $- \Sigma$(bond enthalpies of bonds made)

Exam practice 17

Q1.

The equation shows the reaction of methane with oxygen.



The table shows the bond energies.

Bond	**C–H**	**O=O**	**C=O**	**O–H**
Bond dissociation energy in kJ per mole	412	496	803	463

Calculate the overall energy change for the combustion of one mole of methane.

Energy change = ____________________ kJ mol^{-1}

(3)

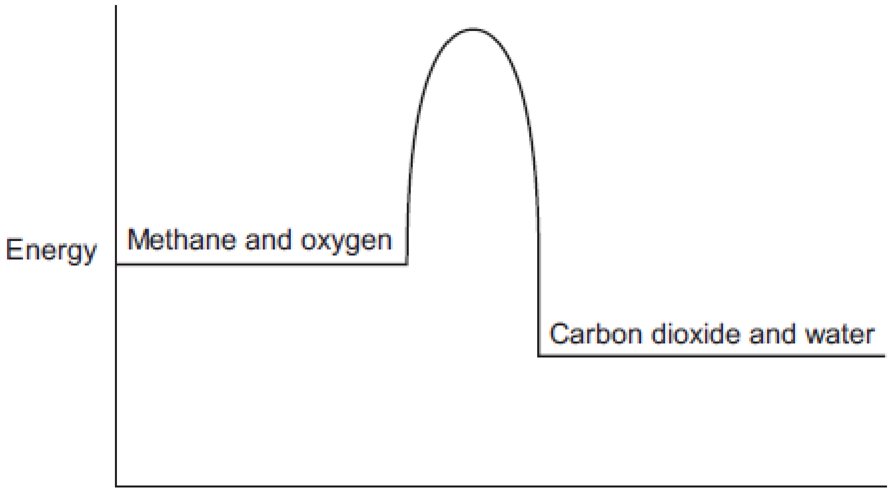
Q2.

(a) Methane burns in oxygen.

(i) The diagram below shows the energy level diagram for the complete combustion of methane.

Draw and label arrows on the diagram to show:

- the activation energy
- the enthalpy change, *ΔH*.



(2)

(ii) Explain why, in terms of the energy involved in bond breaking and bond making, the combustion of methane is exothermic.

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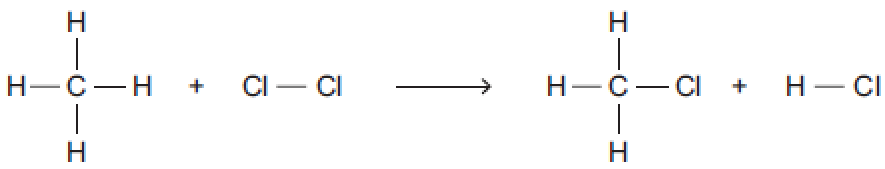
(3)

(b) Methane reacts with chlorine in the presence of sunlight.

The equation for this reaction is:



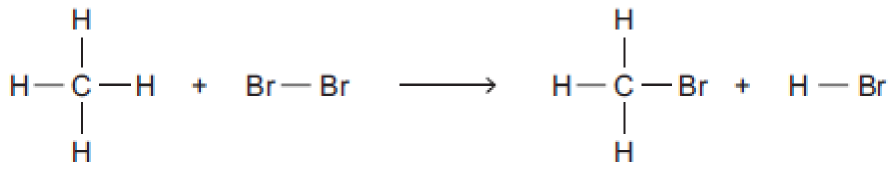
Some bond dissociation energies are given in the table.

Bond	**Bond dissociation energy in kJ per mole**
C–H	413
C–Cl	327
Cl–Cl	243
H–Cl	432

(i) Show that the enthalpy change, ΔH, for this reaction is −103 kJ per mole.

(3)

(ii) Methane also reacts with bromine in the presence of sunlight.



This reaction is less exothermic than the reaction between methane and chlorine.

The enthalpy change, ΔH, is −45 kJ per mole.

What is a possible reason for this?

Tick (✓) **one** box.

CH_3Br has a lower boiling point than CH_3Cl ☐

The C−Br bond is weaker than the C−Cl bond. ☐

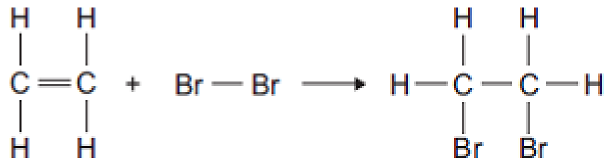
The H−Cl bond is weaker than the H−Br bond. ☐

Chlorine is more reactive than bromine. ☐

(1)

Q3.

(i) The equation for the reaction can be represented as



Bond	Bond dissociation energy in kJ per mole
C—H	413
C═C	614
Br—Br	193
C—C	348
C—Br	276

Use the bond dissociation energies in the table to calculate the enthalpy change (ΔH) for this reaction.

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Enthalpy change (ΔH) = _________ kJ per mole

(3)

(ii) The reaction is exothermic.

Explain why, in terms of bonds broken and bonds formed.

____

____

____

(2)

Q4.

Chlorine reacts with hydrogen to produce hydrogen chloride.

The table shows the values of some bond dissociation energies.

Bond	**H—H**	**Cl—Cl**	**H—Cl**
Dissociation energy in kJ per mole	436	242	431

Use the values in the table to calculate the enthalpy change (ΔH) for the reaction.

$$H_2(g) \quad + \quad Cl_2(g) \longrightarrow 2\ HCl(g)$$

____

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____

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Enthalpy change (ΔH) = ______________ kJ per mole

(3)