Name:

Chemistry

Braintree Sixth Form

Introduction to Inorganic and Physical Chemistry

Terminology

Term	Notes
atom	smallest particle of an element that can exist; consists of a positive nucleus (composed of protons and neutrons) surrounded by negative electrons
ion	an atom with an electrical charge; positive ions are cations and negative ions are anions; molecular ions are molecules with an electrical charge
molecule	a neutral particle made of two or more atoms joined together; the smallest particle of a compound that can exist
element	a pure substance that cannot be chemically broken down into two or more simpler substances; contains just one type of atom (maybe in molecules)
compound	a pure substance that can be chemically broken down into two or more simpler substances; made of just one type of molecule made of different types of atom
relative atomic mass, A_r	the mean mass of the atoms of an element weighted to the relative abundance of its isotope in a.m.u. where 12 a.m.u = the mass of a ^{12}C atom
relative formula mass, M_r	The mass of a molecule or formula unit of a compound in a.m.u. calculated by adding the relative atomic masses of *all* of the atoms within it
formula unit	for some substances, such metals or ionic substances, the empirical formula is taken to represent one unit of the compound, termed a formula unit
molar mass, M_r	the mass in grammes of one mole of molecules/formula units of a substance; molar mass in grammes = relative formula mass in a.m.u.
mole, mol	the unit of the chemical amount of a substance; 1 mole is the same number of particles or formula units as there are atoms in 12g of ^{12}C
(chemical) amount, n	how many particles or formula units there are in a given quantity of a substance; as this number it is mind-bogglingly big, it is counted in moles
concentration, [substance]	the amount of substance per cubic decimetre; for solutions, the amount of substance dissolved per cubic decimetre of solution
cubic decimetre, dm^3	the volume occupied by a cube with an edge of 10 cm (1 dm); 1 dm^3 = 1000 cm^3; 1 dm^3 = 1 litre (**n.b.** 1 cm^3 = 1 ml)
molecular formula	a formula showing the number of atoms of each element in a molecule, e.g. the molecular formula of glucose is $C_6H_{12}O_6$
empirical formula	a formula showing the simplest ratio of atoms of each element in a substance, e.g. the empirical formula of glucose is CH_2O
word equation	e.g. lithium + water → lithium hydroxide + water
chemical equation	e.g. $2Li_{(s)} + 2\ H_2O_{(l)} \rightarrow 2LiOH_{(aq)} + H_{2(g)}$
ionic equation	e.g. $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$ n.b. spectator ions are excluded as they cancel out
molar volume, V_m	the volume of one mole of a gas under certain conditions; at rtp V_m = 24 dm^{-3} n.b. at rtp you can use $n=V \div V_m$; n.b.^{2} if V is in cm^3 then Vm = 24,000 cm^3
room temperature and pressure, rtp	25°C at normal atmospheric pressure; 25°C @ 1 atm; 25°C @ 760 mmHg; **298K @ 101 kPa**
(actual) yield	the mass of product obtained
theoretical yield	the mass of product you should gain calculated from reaction's equation, the chemical amounts of reactants used and the molar mass of the product
percentage yield	percentage yield = $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100\ \%$
atom economy	atom economy = $\frac{\text{molar mass of } \textit{useful} \text{ product}}{\text{total molar mass of all reactants}} \times 100\ \%$
water of crystallisation	the number of moles of water that combine with 1 mole of a formula unit to form crystals e.g. $CuSO_4.5H_2O$
Avogadro constant, N_A	the number of particles in 1 mole of substance; 6.02 x 10^{23}
Stoichiometry	The ratio of reacting species in a balanced equation

Formulae of Ionic Compounds

The formula of an ionic compound is an empirical formula, determined by the charges on the ions that make up the compound, balanced so the overall charge is zero.

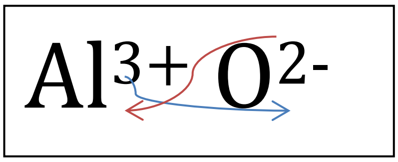
e.g.

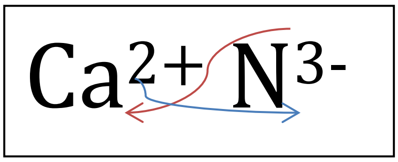
- sodium chloride is composed of Na^+ ions and Cl^- ions.
  The charges balance so we just write the formula as NaCl
- magnesium oxide is composed of Mg^{2+} ions and O^{2-} ions.
  The charges balance so we just write the formula as MgO
- magnesium chloride is composed of Mg^{2+} ions and Cl^- ions.
  To balance we need 2 Cl^- ions per Mg+ ion so we write the formula as $MgCl_2$
- sodium oxide is composed of Na^+ ions and O^{2-} ions.
  To balance we write the formula as Na_2O

Some ions can have charges of 3+ or 3-, when these are combined with ions with a 2- or 2+ charge respectively then we cross the charges to determine the formula.

e.g.

- aluminium oxide is composed of Al^{3+} ions and O^{2-} ions.
  To balance we cross the charges to get Al_2O_3
- calcium nitride is composed of Ca^{2+} ions and N^{3-} ions.
  to balance we cross the charges to get Ca_3N_2





To find the charges on the ions remember these guidelines:

- Metal ions are positive and go with the group they are in, hence group 1 is +, group 2 is 2+ etc.
- Non- metal ions are negative and change their names, gaining a suffix **-ide** and are linked to their group; charge is group no. – 8, thus group 7 is – , group 6 is 2– etc.
- **Hydrogen** ions are H^+ (the only atomic non-metal cation) but **hydride** ions are H^-
- Aside from **ammonium** (NH_4^+), molecular ions are negative and, with the exception of **hydroxide** (OH^-), end in the suffices **–ate** or **–ite**.
- If more than one of a molecular ion in a formula, then the molecular ion should be in brackets with the how many of the ion there written outside the bracket. Otherwise brackets are not needed, thus calcium hydroxide is $Ca(OH)_2$ while potassium hydroxide is KOH.
- The transition metals can have more than one ion and have the a roman numeral after them to indicate their charge, such as Cu^+, copper (I), and Cu^{2+}, copper (II).

Ions to Learn

The following table gives a summary of the ions you should need to know.

Positive ions (cations)			Negative ions (anions)		
Charge	Cation	Symbol	Charge	Anion	Symbol
1+	sodium potassium silver copper(I) hydrogen ammonium	Na^{+} K^{+} Ag^{+} Cu^{+} H^{+} NH_4^{+}	1–	chloride bromide iodide hydroxide nitrate hydrogencarbonate	Cl^{-} Br^{-} I^{-} OH^{-} NO_3^{-} HCO_3^{-}
2+	magnesium calcium zinc copper(II) iron(II)	Mg^{2+} Ca^{2+} Zn^{2+} Cu^{2+} Fe^{2+}	2–	oxide sulfide sulfate sulfite carbonate	O^{2-} S^{2-} SO_4^{2-} SO_3^{2-} CO_3^{2-}
3+	aluminium iron(III)	Al^{3+} Fe^{3+}	3–	nitride phosphate	N^{3-} PO_4^{3-}

Practice 1

Write the formulae for the following: Mark: ________/15

1. Lithium Fluoride ....
2. Magnesium Bromide ....
3. Iron (II) Sulfide ....
4. Rubidium Carbonate ....
5. Silver Nitrate ....
6. Ammonium Hydroxide ....
7. Zinc Hydroxide ....
8. Ammonium Sulfite ....
9. Iron (III) Phosphate ....
10. Iron (II) Phosphate ....
11. Aluminium Sulfate ....
12. Magnesium Hydroxide ....
13. Lithium Hydride ....
14. Aluminium Sulfate ....
15. Copper (II) Nitrate ....

Formulae of Covalent Molecules

Aside from metals (which are written with a formula unit of a single atom) and the group 0 (Noble Gases, which are monatomic), elements exist in a molecular form. Most commonly, elements are **diatomic** molecules, and so we have N_2, O_2, S_2, F_2, Cl_2, Br_2 and I_2

Some elements exist in more than one molecule. These different forms are called **allotropes** and examples are:
Sulphur: most commonly S_2 but also S
Oxygen: O_2 and less commonly O_3 (ozone)
Carbon: two giant molecular forms, graphite and diamond, formula written as C (the formula unit) and fullerene, C_{60}

Here is a list of the common covalent compounds you will come across:

Water	H_2O	Carbon Dioxide	CO_2	Carbon Monoxide	CO
Ammonia	NH_3	Silicon Dioxide	SiO_2	Glucose	$C_6H_{12}O_6$
Hydrogen Chloride	HCl	Sulfur Dioxide	SO_2	Sulfur Trioxide	SO_3
Hydrogen Peroxide	H_2O_2	Nitrogen Dioxide	NO_2	Nitrogen Monoxide	NO
Nitrous Oxide	N_2O	Dinitrogen Tetroxide	N_2O_4	Phosphorus Trichloride	PCl_3

Organic compounds are molecular and their naming is covered in the organic chemistry unit. However, the ones you should know from earlier study are:

Alkanes – general formula $C_nH_{(2n+2)}$
Alkenes – general formula C_nH_{2n}
Alcohols – general formula $C_nH_{(2n+1)}OH$
Carboxylic Acids – general formula $C_{(n-1)}H_{(2n-1)}COOH$

n	1	2	3	4	5	6	7	8
Prefix alkanes, alkenes	meth-	eth-	prop-	but-	pent-	hex-	hept-	oct-
Prefix everything else	methan-	ethan-	propan-	butan-	pentan-	hexan-	heptan-	octan-

n	9	10	11	12	15	17	20
Prefix alkanes, alkenes	non-	dec-	undec-	dodec-	pentdec-	heptadec-	icos-
Prefix everything else	nonan-	decan-	undecan-	dodecan-	pentdecan-	heptadecan-	icosan-

Combustion of Elements

When elements burn, they react with oxygen and so are oxidised. The general equation is:

- "element" + oxygen → "element" oxide

There are exceptions to this, mainly due to the naming of the oxide, e.g.

- hydrogen + oxygen → water
- carbon + oxygen → carbon dioxide
- sulphur + oxygen → sulphur trioxide

Combustion of Common Organic Compounds

Alkanes, alkenes, alcohols, carboxylic acids and sugars all combust to the same general equation and can be considered fuels. The general equation is:

- "fuel" + oxygen → carbon dioxide + water

Practice II

Write **word equations** for the combustion of the following substances

1. Magnesium ....
2. Aluminium ....
3. Carbon ....
4. Rubidium ....
5. Glucose ....
6. Pentane ....
7. Ethene ....
8. Sulfur ....
9. Calcium ....
10. Methanoic Acid ....
11. Ethanol ....
12. Sucrose ....
13. Potassium ....
14. Strontium ....
15. Hydrogen ....

Mark: ________/15

Practice III

Write the formulae of the following oxides:

1. water ....
2. carbon monoxide ....
3. silicon dioxide ....
4. dinitrogen monoxide ....
5. sodium oxide ....
6. nitrogen dioxide ....
7. dinitrogen tetroxide ....
8. sulphur trioxide ....
9. lead (II) oxide ....
10. iron (III) oxide ....
11. copper (I) oxide ....
12. carbon dioxide ....
13. nitrous oxide ....
14. copper (II) oxide ....
15. nitrogen monoxide ....

Mark: ________/15

Yields

The mass of product made by a process is called its yield. There are three types of yield:

- The maximum mass of product a process can make is called the theoretical yield.
- The actual mass of product made by the process is called the actual yield.
- The percentage yield is the actual yield expressed as a percentage of the theoretical yield:

$$Percentage\ Yield = \frac{Actual\ Yield}{Theoretical\ Yield} \times 100\ \%$$

Practice IV

This is the procedure to extract sodium chloride from an impure sample. Complete the table and work out the yield.

Procedure

1. Collect a beaker (250 ml) and accurately measure its mass.
2. Pour your sample of impure sodium chloride into the beaker.
3. Accurately measure the mass of the beaker and impure sodium chloride.
4. Measure out 50 cm^3 of distilled water (100 ml measuring cylinder) from the plastic aspirator.
5. Pour the water into your beaker and bring the water to the boil (Bunsen burner) while stirring it constantly with a glass rod.
6. Once the water is boiling, turn off the heat and continue to stir for a further five minutes.
7. Collect a porcelain evaporating basin and accurately measure its mass.
8. Carefully filter the hot liquid into the evaporating basin.
9. Boil away the water to leave solid sodium chloride behind.
10. Accurately measure the mass of the beaker and sodium chloride.
11. Return the evaporating basin to the heat for a further minute and measure its mass once more.
12. Repeat step 11 until the mass of the basin and sodium chloride is consistent.

Results Mark: ____________/10

Reading	Measurement	Unit
Mass of beaker	87.18	g
Mass of beaker + impure sodium chloride	89.23	g
∴ Mass of impure sodium chloride		g
Mass of evaporating basin	113.56	g
Consistent mass of evaporating basin and pure sodium chloride	115.20	g
∴ Mass of pure sodium chloride		g

100 g of the impure sodium chloride should yield 42 g of pure sodium chloride.

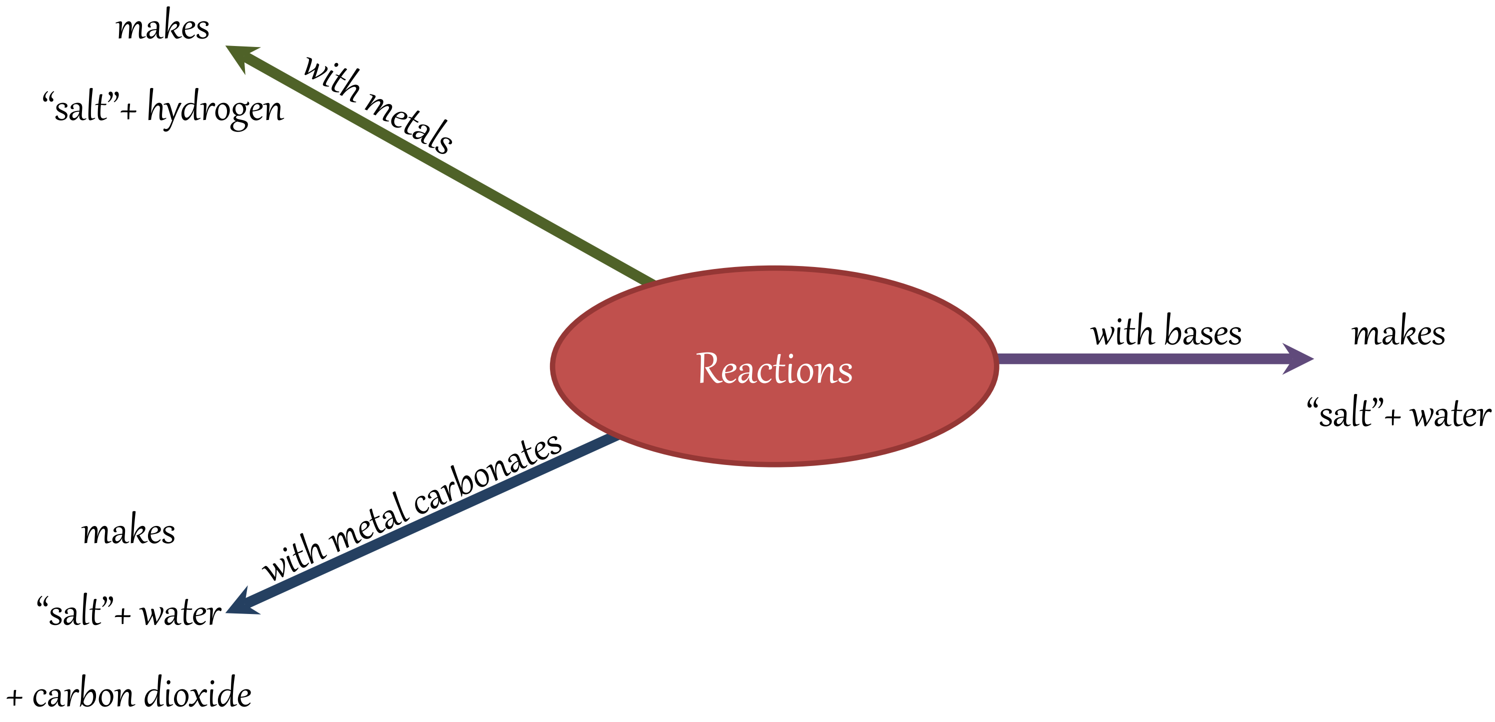
Theoretical yield for this process for the mass of impure sodium chloride used = __________ g

Actual yield = __________ g

∴ Percentage yield = __________ %

Reactions of Acids

There are three general reactions of acids:



These can be written as:

- "acid" + "metal" →"salt" + hydrogen
- "acid" + "base" →"salt" + water
- "acid" + "metal carbonate" →"salt" + water + carbon dioxide

Salt is another name for an ionic compound, specifically one made by the above reactions.
Bases are metal oxides or metal hydroxides. There is a subgroup of bases, which are those that are soluble in water, called **alkalis**.
Ammonia is worth mentioning here, as it readily reacts with water forming ammonium hydroxide, which is an alkali. The reaction of ammonium hydroxide with an acid is:

- ammonium hydroxide + "acid" → "ammonium salt" + water

However, as ammonia bubbled through an acid will react with the water the acid is dissolved in, We get a two stage reaction:

- ammonia + water → ammonium hydroxide + "acid" → "ammonium salt" + water

The water cancels out , when the reaction is simplified, and so we get:

- ammonia + "acid" → "ammonium salt"

Just as ammonium hydroxide is considered to be an alkali, ammonium carbonate is treated as if it were a metal carbonate. Also, metal hydrogencarbonates (sometimes called bicarbonates) react in this way with acids.

Reactions of Acids (continued)

As ionic compounds, salts are composed of a cation and an anion. Obviously the cation in the salt is the metal, either that which reacts with the acid or forms part of the base or carbonate that reacts with the acid (or ammonium if appropriate). Similarly, the anion is the anion in the acid. Here are some acids and their anions (the most common acids you will come across are **emboldened**):

Acid	Anion	Symbol
hydrofluoric	fluoride	F^-
hydrochloric	**chloride**	Cl^-
hydrobromic	bromide	Br^-
hydroiodic	iodide	I^-
nitric	**nitrate**	NO_3^-
methanoic	methanoate	$HCOO^-$
ethanoic	**ethanoate**	CH_3COO^-
etc.		
sulfuric	sulfate	SO_4^{2-}
sulfurous	sulfite	SO_3^{2-}
phosphoric	**phosphate**	PO_4^{3-}

Practice V

Complete the **word equations** for following reactions: Mark: ____________/15

1. hydrochloric acid + sodium → ....
2. nitric acid + ammonium hydroxide → ....
3. ethanoic acid + lithium oxide → ....
4. sulfuric acid + calcium carbonate → ....
5. phosphoric acid + iron (III) oxide → ....
6. hydrochloric acid + zinc → ....
7. nitric acid + calcium hydroxide → ....
8. ethanoic acid + magnesium oxide → ....
9. sulfuric acid + iron (III) carbonate → ....
10. phosphoric acid + copper (II) oxide → ....
11. hydrochloric acid + iron (III) oxide → ....
12. nitric acid + potassium hydroxide → ....
13. ethanoic acid + lead (II) oxide → ....
14. sulfuric acid + ammonium carbonate → ....
15. phosphoric acid + silver oxide → ....

Practice VI

Write **balanced chemical equations** for following reactions: Mark: ____________/15

1. hydrochloric acid + sodium :

2. nitric acid + ammonium hydroxide :

3. ethanoic acid + lithium oxide :

4. sulfuric acid + calcium carbonate :

5. phosphoric acid + iron (III) oxide :

6. hydrochloric acid + zinc :

7. nitric acid + calcium hydroxide :

8. ethanoic acid + magnesium oxide :

9. sulfuric acid + iron (III) carbonate :

10. phosphoric acid + copper (II) oxide :

11. hydrochloric acid + iron (III) oxide :

12. nitric acid + potassium hydroxide :

13. ethanoic acid + lead (II) oxide :

Practice VI (continued)

14. sulfuric acid + ammonium carbonate :

15. phosphoric acid + silver oxide :

Stoichiometry

Consider the equation for the combustion of hydrogen:

hydrogen + oxygen → water

$2H_2 + O_2 \rightarrow 2H_2O$

The large numbers at the start of the equations mean that two molecules of hydrogen react with one molecule of oxygen to make two molecules of water. This ratio of reacting species is called **stoichiometry** and it is very useful. Sometimes this is called the molar ratio.

Practice VII

Use stoichiometry to find out how many molecules of the second substance react with six molecules of the acid named. Mark: ____________/15

1. hydrochloric acid + sodium →
2. nitric acid + ammonium hydroxide →
3. ethanoic acid + lithium oxide →
4. sulfuric acid + calcium carbonate →
5. phosphoric acid + iron (III) oxide →
6. hydrochloric acid + zinc →
7. nitric acid + calcium hydroxide →
8. ethanoic acid + magnesium oxide →
9. sulfuric acid + iron (III) carbonate →
10. phosphoric acid + copper (II) oxide →
11. hydrochloric acid + iron (III) oxide →
12. nitric acid + potassium hydroxide →
13. ethanoic acid + lead (II) oxide →
14. sulfuric acid + ammonium carbonate →
15. phosphoric acid + silver oxide →

The Periodic Table of the Elements

Key

relative atomic mass
symbol
name
atomic (proton) number

1	2											3	4	5	6	7	0
(1)	(2)	(3)	(4)	(5)	(6)	(7)	(8)	(9)	(10)	(11)	(12)	(13)	(14)	(15)	(16)	(17)	(18)
							1.0 **H** hydrogen 1										4.0 **He** helium 2
6.9 **Li** lithium 3	9.0 **Be** beryllium 4											10.8 **B** boron 5	12.0 **C** carbon 6	14.0 **N** nitrogen 7	16.0 **O** oxygen 8	19.0 **F** fluorine 9	20.2 **Ne** neon 10
23.0 **Na** sodium 11	24.3 **Mg** magnesium 12											27.0 **Al** aluminium 13	28.1 **Si** silicon 14	31.0 **P** phosphorus 15	32.1 **S** sulfur 16	35.5 **Cl** chlorine 17	39.9 **Ar** argon 18
39.1 **K** potassium 19	40.1 **Ca** calcium 20	45.0 **Sc** scandium 21	47.9 **Ti** titanium 22	50.9 **V** vanadium 23	52.0 **Cr** chromium 24	54.9 **Mn** manganese 25	55.8 **Fe** iron 26	58.9 **Co** cobalt 27	58.7 **Ni** nickel 28	63.5 **Cu** copper 29	65.4 **Zn** zinc 30	69.7 **Ga** gallium 31	72.6 **Ge** germanium 32	74.9 **As** arsenic 33	79.0 **Se** selenium 34	79.9 **Br** bromine 35	83.8 **Kr** krypton 36
85.5 **Rb** rubidium 37	87.6 **Sr** strontium 38	88.9 **Y** yttrium 39	91.2 **Zr** zirconium 40	92.9 **Nb** niobium 41	96.0 **Mo** molybdenum 42	[98] **Tc** technetium 43	101.1 **Ru** ruthenium 44	102.9 **Rh** rhodium 45	106.4 **Pd** palladium 46	107.9 **Ag** silver 47	112.4 **Cd** cadmium 48	114.8 **In** indium 49	118.7 **Sn** tin 50	121.8 **Sb** antimony 51	127.6 **Te** tellurium 52	126.9 **I** iodine 53	131.3 **Xe** xenon 54
132.9 **Cs** caesium 55	137.3 **Ba** barium 56	138.9 **La** * lanthanum 57	178.5 **Hf** hafnium 72	180.9 **Ta** tantalum 73	183.8 **W** tungsten 74	186.2 **Re** rhenium 75	190.2 **Os** osmium 76	192.2 **Ir** iridium 77	195.1 **Pt** platinum 78	197.0 **Au** gold 79	200.6 **Hg** mercury 80	204.4 **Tl** thallium 81	207.2 **Pb** lead 82	209.0 **Bi** bismuth 83	[209] **Po** polonium 84	[210] **At** astatine 85	[222] **Rn** radon 86
[223] **Fr** francium 87	[226] **Ra** radium 88	[227] **Ac** † actinium 89	[267] **Rf** rutherfordium 104	[268] **Db** dubnium 105	[271] **Sg** seaborgium 106	[272] **Bh** bohrium 107	[270] **Hs** hassium 108	[276] **Mt** meitnerium 109	[281] **Ds** darmstadtium 110	[280] **Rg** roentgenium 111	Elements with atomic numbers 112-116 have been reported but not fully authenticated						

* 58 – 71 Lanthanides	140.1 **Ce** cerium 58	140.9 **Pr** praseodymium 59	144.2 **Nd** neodymium 60	[145] **Pm** promethium 61	150.4 **Sm** samarium 62	152.0 **Eu** europium 63	157.3 **Gd** gadolinium 64	158.9 **Tb** terbium 65	162.5 **Dy** dysprosium 66	164.9 **Ho** holmium 67	167.3 **Er** erbium 68	168.9 **Tm** thulium 69	173.1 **Yb** ytterbium 70	175.0 **Lu** lutetium 71
† 90 – 103 Actinides	232.0 **Th** thorium 90	231.0 **Pa** protactinium 91	238.0 **U** uranium 92	[237] **Np** neptunium 93	[244] **Pu** plutonium 94	[243] **Am** americium 95	[247] **Cm** curium 96	[247] **Bk** berkelium 97	[251] **Cf** californium 98	[252] **Es** einsteinium 99	[257] **Fm** fermium 100	[258] **Md** mendelevium 101	[259] **No** nobelium 102	[262] **Lr** lawrencium 103

Transition from GCSE to A Level

Moving from GCSE Science to A Level can be a daunting leap. You'll be expected to remember a lot more facts, equations, and definitions, and you will need to learn new maths skills and develop confidence in applying what you already know to unfamiliar situations.

This worksheet aims to give you a head start by helping you:

- to pre-learn some useful knowledge from the first chapters of your A Level course
- understand and practise of some of the maths skills you'll need.

Learning objectives

After completing the worksheet you should be able to:

- define practical science key terms
- recall the answers to the retrieval questions
- perform maths skills including:
  - converting between units and standard form and decimals
  - balancing chemical equations
  - rearranging equations
  - calculating moles and masses
  - calculating percentage yield and percentage error
  - interpreting graphs of reactions.

Retrieval questions

You need to be confident about the definitions of terms that describe measurements and results in A Level Chemistry.

Learn the answers to the questions below then cover the answers column with a piece of paper and write as many answers as you can. Check and repeat.

Practical science key terms

When is a measurement valid?	when it measures what it is supposed to be measuring
When is a result accurate?	when it is close to the true value
What are precise results?	when repeat measurements are consistent/agree closely with each other
What is repeatability?	how precise repeated measurements are when they are taken by the *same* person, using the *same* equipment, under the *same* conditions
What is reproducibility?	how precise repeated measurements are when they are taken by *different* people, using *different* equipment
What is the uncertainty of a measurement?	the interval within which the true value is expected to lie
Define measurement error	the difference between a measured value and the true value
What type of error is caused by results varying around the true value in an unpredictable way?	random error
What is a systematic error?	a consistent difference between the measured values and true values
What does zero error mean?	a measuring instrument gives a false reading when the true value should be zero
Which variable is changed or selected by the investigator?	independent variable
What is a dependent variable?	a variable that is measured every time the independent variable is changed
Define a fair test	a test in which only the independent variable is allowed to affect the dependent variable
What are control variables?	variables that should be kept constant to avoid them affecting the dependent variable

Atomic structure

Learn the answers to the questions below then cover the answers column with a piece of paper and write as many answers as you can. Check and repeat.

What does an atom consist of?	a nucleus containing protons and neutrons, surrounded by electrons
What are the relative masses of a proton, neutron, and electron?	1, 1, and $\frac{1}{1840}$ respectively
What are the relative charges of a proton, neutron, and electron?	+1, 0, and -1 respectively
How do the number of protons and electrons differ in an atom?	they are the same because atoms have neutral charge
What force holds an atomic nucleus together?	strong nuclear force
What is the atomic number of an element?	the number of protons in the nucleus of a single atom of an element
What is the mass number of an element?	number of protons + number of neutrons
What is an isotope?	an atom with the same number of protons but different number of neutrons
What is an ion?	an atom, or group of atoms, with a charge
What is the function of a mass spectrometer?	it accurately determines the mass and abundance of separate atoms or molecules, to help us identify them
What is a mass spectrum?	the output from a mass spectrometer that shows the different isotopes that make up an element
What is the total number of electrons that each electron shell (main energy level) can contain?	$2n^2$ electrons, where n is the number of the shell
How many electrons can the first three electron shells hold each?	2 electrons (first shell), 8 electrons (second shell), 18 electrons (third shell)
What are the first four electron sub-shells (orbitals) called?	s, p, d, and f (in order)
How many electrons can each orbital hold?	a maximum of 2 electrons
Define the term ionisation energy, and give its unit	the energy it takes to remove a mole of electrons from a mole of atoms in the gaseous state, unit = kJ mol^{-1}
What is the equation for relative atomic mass (A_r)?	$\text{relative atomic mass} = \frac{\text{average mass of 1 atom}}{\frac{1}{12}^{th}\ \text{mass of 1 atom of}\ ^{12}C}$
What is the equation for relative molecular mass (M_r)?	$\text{relative molecular mass} = \frac{\text{average mass of 1 molecule}}{\frac{1}{12}^{th}\ \text{mass of 1 atom of}\ ^{12}C}$

Maths skills

1 Core mathematical skills

A practical chemist must be proficient in standard form, significant figures, decimal places, SI units, and unit conversion.

1.1 Standard form

In science, very large and very small numbers are usually written in standard form. Standard form is writing a number in the format A × 10^x where A is a number from 1 to 10 and x is the number of places you move the decimal place.

For example, to express a large number such as 50 000 mol dm^{-3} in standard form, A = 5 and x = 4 as there are four numbers after the initial 5.

Therefore, it would be written as 5×10^4 mol dm^{-3}.

To give a small number such as 0.000 02 Nm^2 in standard form, A = 2 and there are five numbers before it so x = −5.

So it is written as 2×10^{-5} Nm^2.

Practice questions

1 Change the following values to standard form.

a boiling point of sodium chloride: 1413 °C

b largest nanoparticles: 0.0 001×10^{-3} m

c number of atoms in 1 mol of water: 1806×10^{21}

2 Change the following values to ordinary numbers.

a 5.5×10^{-6} **b** 2.9×10^2 **c** 1.115×10^4 **d** 1.412×10^{-3} **e** 7.2×10^1

1.2 Significant figures and decimal places

In chemistry, you are often asked to express numbers to either three or four significant figures. The word significant means to ‘have meaning’. A number that is expressed in significant figures will only have digits that are important to the number’s precision.

It is important to record your data and your answers to calculations to a reasonable number of significant figures. Too many and your answer is claiming an accuracy that it does not have, too few and you are not showing the precision and care required in scientific analysis.

For example, 6.9301 becomes 6.93 if written to three significant figures.

Likewise, 0.000 434 56 is 0.000 435 to three significant figures.

Notice that the zeros before the figure are *not* significant – they just show you how large the number is by the position of the decimal point. Here, a 5 follows the last significant digit, so just as with decimals, it must be rounded up.

Any zeros between the other significant figures are significant. For example, 0.003 018 is 0.003 02 to three significant figures.

Sometimes numbers are expressed to a number of decimal places. The decimal point is a place holder and the number of digits afterwards is the number of decimal places.

For example, the mathematical number pi is 3 to zero decimal places, 3.1 to one decimal place, 3.14 to two decimal places, and 3.142 to three decimal places.

Practice questions

3 Give the following values in the stated number of significant figures (s.f.).

a 36.937 (3 s.f.) **b** 258 (2 s.f.) **c** 0.043 19 (2 s.f.) **d** 7 999 032 (1 s.f.)

4 Use the equation:

number of molecules = number of moles × 6.02 × 10^{23} molecules per mole

to calculate the number of molecules in 0.5 moles of oxygen. Write your answer in standard form to 3 s.f.

5 Give the following values in the stated number of decimal places (d.p.).

a 4.763 (1 d.p.) **b** 0.543 (2 d.p.) **c** 1.005 (2 d.p.) **d** 1.9996 (3 d.p.)

1.3 Converting units

Units are defined so that, for example, every scientist who measures a mass in kilograms uses the same size for the kilogram and gets the same value for the mass. Scientific measurement depends on standard units – most are *Système International* (SI) units.

If you convert between units and round numbers properly it allows quoted measurements to be understood within the scale of the observations.

Multiplication factor	Prefix	Symbol
10^{9}	giga	G
10^{6}	mega	M
10^{3}	kilo	k
10^{-2}	centi	c
10^{-3}	milli	m
10^{-6}	micro	μ
10^{-9}	nano	n

Unit conversions are common. For instance, you could be converting an enthalpy change of 488 889 J mol^{-1} into kJ mol^{-1}. A kilo is 10^{3} so you need to divide by this number or move the decimal point three places to the left.

488 889 ÷ 10^{3} kJ mol^{-1} = 488.889 kJ mol^{-1}

Converting from mJ mol^{-1} to kJ mol^{-1}, you need to go from 10^{3} to 10^{-3}, or move the decimal point six places to the left.

333 mJ mol^{-1} is 0.000 333 kJ mol^{-1}

If you want to convert from 333 mJ mol^{-1} to nJ mol^{-1}, you would have to go from 10^{-9} to 10^{-3}, or move the decimal point six places to the right.

333 mJ mol^{-1} is 333 000 000 nJ mol^{-1}

Practice questions

6 Calculate the following unit conversions.

a 300 µm to m

b 5 MJ to mJ

c 10 GW to kW

2 Balancing chemical equations

2.1 Conservation of mass

When new substances are made during chemical reactions, atoms are not created or destroyed – they just become rearranged in new ways. So, there is always the same number of each type of atom before and after the reaction, and the total mass before the reaction is the same as the total mass after the reaction. This is known as the conservation of mass.

You need to be able to use the principle of conservation of mass to write formulae, and balanced chemical equations and half equations.

2.2 Balancing an equation

The equation below shows the correct formulae but it is not balanced.

$H_2 + O_2 \rightarrow H_2O$

While there are two hydrogen atoms on both sides of the equation, there is only one oxygen atom on the right-hand side of the equation against two oxygen atoms on the left-hand side. Therefore, a two must be placed before the H_2O.

$H_2 + O_2 \rightarrow 2H_2O$

Now the oxygen atoms are balanced but the hydrogen atoms are no longer balanced. A two must be placed in front of the H_2.

$2H_2 + O_2 \rightarrow 2H_2O$

The number of hydrogen and oxygen atoms is the same on both sides, so the equation is balanced.

Practice questions

1 Balance the following equations.

a $C + O_2 \rightarrow CO$

b $N_2 + H_2 \rightarrow NH_3$

c $C_2H_4 + O_2 \rightarrow H_2O + CO_2$

2.3 Balancing an equation with fractions

To balance the equation below:

$C_2H_6 + O_2 \rightarrow CO_2 + H_2O$

- Place a two before the CO_2 to balance the carbon atoms.
- Place a three in front of the H_2O to balance the hydrogen atoms.

  $C_2H_6 + O_2 \rightarrow 2CO_2 + 3H_2O$

There are now four oxygen atoms in the carbon dioxide molecules plus three oxygen atoms in the water molecules, giving a total of seven oxygen atoms on the product side.

- To balance the equation, place three and a half in front of the O_2.

  $C_2H_6 + 3\frac{1}{2}O_2 \rightarrow 2CO_2 + 3H_2O$

- Finally, multiply the equation by 2 to get whole numbers.

  $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$

Practice questions

2 Balance the equations below.

a $C_6H_{14} + O_2 \rightarrow CO_2 + H_2O$

b $NH_2CH_2COOH + O_2 \rightarrow CO_2 + H_2O + N_2$

2.4 Balancing an equation with brackets

$Ca(OH)_2 + HCl \rightarrow CaCl_2 + H_2O$

Here the brackets around the hydroxide (OH^-) group show that the $Ca(OH)_2$ unit contains one calcium atom, two oxygen atoms, and two hydrogen atoms.

To balance the equation, place a two before the HCl and another before the H_2O.

$Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O$

Practice questions

3 Balance the equations below.

a $Mg(OH)_2 + HNO_3 \rightarrow Mg(NO_3)_2 + H_2O$

b $Fe(NO_3)_2 + Na_3PO_4 \rightarrow Fe_3(PO_4)_2 + NaNO_3$

3 Rearranging equations and calculating concentrations

3.1 Rearranging equations

In chemistry, you sometimes need to rearrange an equation to find the desired values.

For example, you may know the amount of a substance (*n*) and the mass of it you have (*m*), and need to find its molar mass (*M*).

The amount of substance (*n*) is equal to the mass you have (*m*) divided by the molar mass (*M*):

$$n = \frac{m}{M}$$

You need to rearrange the equation to make the molar mass (*M*) the subject.

Multiply both sides by the molar mass (*M*):

$$M \times n = m$$

Then divide both sides by the amount of substance (*n*):

$$m = \frac{m}{N}$$

Practice questions

1 Rearrange the equation $c = \frac{n}{V}$ to make:

a *n* the subject of the equation

b *V* the subject of the equation.

2 Rearrange the equation $PV = nRT$ to make:

a *n* the subject of the equation

b *T* the subject of the equation.

3.2 Calculating concentration

The concentration of a solution (a solute dissolved in a solvent) is a way of saying how much solute, in moles, is dissolved in 1 dm^3 or 1 litre of solution.

Concentration is usually measured using units of mol dm^{-3}. (It can also be measured in g dm^3.)

The concentration of the amount of substance dissolved in a given volume of a solution is given by the equation:

$$c = \frac{n}{V}$$

where *n* is the amount of substance in moles, *c* is the concentration, and *V* is the volume in dm^3.

The equation can be rearranged to calculate:

- the amount of substance *n*, in moles, from a known volume and concentration of solution
- the volume *V* of a solution from a known amount of substance, in moles, and the concentration of the solution.

Practice questions

3 Calculate the concentration, in mol dm^{-3}, of a solution formed when 0.2 moles of a solute is dissolved in 50 cm^3 of solution.

4 Calculate the concentration, in mol dm^{-3}, of a solution formed when 0.05 moles of a solute is dissolved in 2.0 dm^3 of solution.

5 Calculate the number of moles of NaOH in an aqueous solution of 36 cm^3 of 0.1 mol dm^{-3}.

4 Molar calculations

4.1 Calculating masses and gas volumes

The balanced equation for a reaction shows how many moles of each reactant and product are involved in a chemical reaction.

If the amount, in moles, of one of the reactants or products is known, the number of moles of any other reactants or products can be calculated.

The number of moles (*n*), the mass of the substance (*m*), and the molar mass (*M*) are linked by:

$$n = \frac{m}{M}$$

Note: The molar mass of a substance is the mass per mole of the substance. For $CaCO_3$, for example, the atomic mass of calcium is 40.1, carbon is 12, and oxygen is 16. So the molar mass of $CaCO_3$ is:

40.1 + 12 + (16 × 3) = 100.1. The units are g mol^{-1}.

Look at this worked example. A student heated 2.50 g of calcium carbonate, which decomposed as shown in the equation:

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

The molar mass of calcium carbonate is 100.1 g mol^{-1}.

a Calculate the amount, in moles, of calcium carbonate that decomposes.

$$n = \frac{m}{M} = 2.50/100.1 = 0.025\ \text{mol}$$

b Calculate the amount, in moles, of carbon dioxide that forms.

From the balanced equation, the number of moles of calcium carbonate = number of moles of carbon dioxide = 0.025 mol

Practice questions

1 In a reaction, 0.486 g of magnesium was added to oxygen to produce magnesium oxide.

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

a Calculate the amount, in moles, of magnesium that reacted.

b Calculate the amount, in moles, of magnesium oxide made.

c Calculate the mass, in grams, of magnesium oxide made.

2 Oscar heated 4.25 g of sodium nitrate. The equation for the decomposition of sodium nitrate is:

$2NaNO_3(s) \rightarrow 2NaNO_2(s) + O_2(g)$

a Calculate the amount, in moles, of sodium nitrate that reacted.

b Calculate the amount, in moles, of oxygen made.

3 0.500 kg of magnesium carbonate decomposes on heating to form magnesium oxide and carbon dioxide. Give your answers to 3 significant figures.

$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$

a Calculate the amount, in moles, of magnesium carbonate used.

b Calculate the amount, in moles, of carbon dioxide produced.

5 Percentage yields and percentage errors

5.1 Calculating percentage yield

Chemists often find that an experiment makes a smaller amount of product than expected. They can predict the amount of product made in a reaction by calculating the percentage yield.

The percentage yield links the actual amount of product made, in moles, and the theoretical yield, in moles:

$$\text{percentage yield} = \frac{\text{actual amount (in moles) of product}}{\text{theoretical amount (in moles) of product}} \times 100$$

Look at this worked example. A student added ethanol to propanoic acid to make the ester, ethyl propanoate, and water.

$C_2H_5OH + C_2H_5COOH \rightarrow C_2H_5COOC_2H_5 + H_2O$

The experiment has a theoretical yield of 5.00 g.

The actual yield is 4.50 g.

The molar mass of $C_2H_5COOC_2H_5$ = 102.0 g mol^{-1}

Calculate the percentage yield of the reaction.

Actual amount of ethyl propanoate: $n = \frac{m}{M}$ = 4.5/102 = 0.0441 mol

Theoretical amount of ethyl propanoate: $n = \frac{m}{M}$ = 5.0/102 = 0.0490 mol

percentage yield = (0.0441/0.0490) × 100% = 90%

Practice questions

1 Calculate the percentage yield of a reaction with a theoretical yield of 4.75 moles of product and an actual yield of 3.19 moles of product. Give your answer to 3 significant figures.

2 Calculate the percentage yield of a reaction with a theoretical yield of 12.00 moles of product and an actual yield of 6.25 moles of product. Give your answer to 3 significant figures.

5.2 Calculating percentage error in apparatus

The percentage error of a measurement is calculated from the maximum error for the piece of apparatus being used and the value measured:

$$\text{percentage error} = \frac{\text{maximum error}}{\text{measured value}} \times 100\%$$

Look at this worked example. In an experiment to measure temperature changes, an excess of zinc powder was added to 50 cm^3 of copper(II) sulfate solution to produce zinc sulfate and copper.

$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$

The measuring cylinder used to measure the copper(II) sulfate solution has a maximum error of ±2 cm^3.

a Calculate the percentage error.

percentage error = (2/50) × 100% = 4%

b A thermometer has a maximum error of ±0.05 °C.

Calculate the percentage error when the thermometer is used to record a temperature rise of 3.9 °C. Give your answer to 3 significant figures.

percentage error = (2 × 0.05)/3.9 × 100% = 2.56%

(Notice that two measurements of temperature are required to calculate the temperature change so the maximum error is doubled.)

Practice questions

3 A gas syringe has a maximum error of ±0.5 cm^3. Calculate the maximum percentage error when recording these values. Give your answers to 3 significant figures.

a 21.0 cm^3 **b** 43.0 cm^3

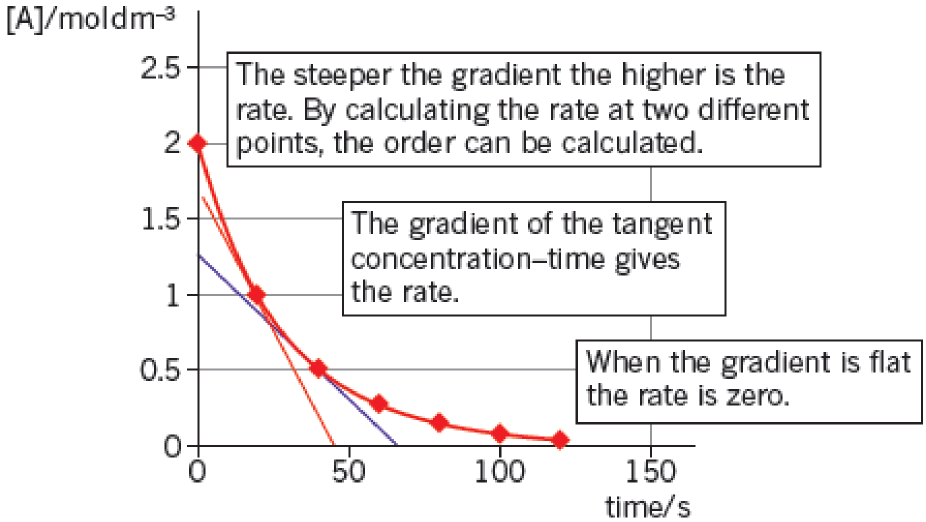
4 A thermometer has a maximum error of ±0.5 °C. Calculate the maximum percentage error when recording these temperature rises. Give your answers to 3 significant figures.

a 12.0 °C **b** 37.6 °C

6 Graphs and tangents

6.1 Deducing reaction rates

To investigate the reaction rate during a reaction, you can measure the volume of the product formed, such as a gas, or the colour change to work out the concentration of a reactant during the experiment. By measuring this concentration at repeated intervals, you can plot a concentration–time graph.



Note: When a chemical is listed in square brackets, it just means 'the concentration of' that chemical. For example, $[O_2]$ is just shorthand for the concentration of oxygen molecules.

By measuring the gradient (slope) of the graph, you can calculate the rate of the reaction. In the graph above, you can see that the gradient changes as the graph is a curve. If you want to know the rate of reaction when the graph is curved, you need to determine the gradient of the curve. So, you need to plot a tangent.

The tangent is the straight line that just touches the curve. The gradient of the tangent is the gradient of the curve at the point where it touches the curve.

Looking at the graph above. When the concentration of A has halved to 1.0 mol dm^{-3}, the tangent intercepts the *y*-axis at 1.75 and the *x*-axis at 48.

The gradient is $\frac{-1.75}{48} = -0.0365$ (3 s.f.).

So the rate is 0.0365 mol dm^{-3} s^{-1}.

Practice questions

1 Using the graph above, calculate the rate of reaction when the concentration of A halves again to 0.5 mol dm^{-3}.

6.2 Deducing the half-life of a reactant

In chemistry, half-life can also be used to describe the decrease in concentration of a reactant in a reaction. In other words, the half-life of a reactant is the time taken for the concentration of the reactant to fall by half.

Practice questions

2 The table below shows the change in concentration of bromine during the course of a reaction.

Time / s	$[Br_2]$ / mol dm^{-3}
0	0.0100
60	0.0090
120	0.0066
180	0.0053
240	0.0044
360	0.0028

a Plot a concentration–time graph for the data in the table.

b Calculate the rate of decrease of Br_2 concentration by drawing tangents.

c Find the half-life at two points and deduce the order of the reaction.

Answers to maths skills practice questions

1 Core mathematics

1 **a** 1.413×10^3 °C **b** 1.0×10^{-7} m
c 1.806×10^{21} atoms

2 **a** 0.000 0055 **b** 290
c 11150 **d** 0.001 412
e 72

3 **a** 36.9 **b** 260
c 0.043 **d** 8 000 000

4 Number of molecules = 0.5 moles × $6.022 \times 10^{23} = 3.011 \times 10^{23} = 3.01 \times 10^{23}$

5 **a** 4.8 **b** 0.54
c 1.01 **d** 2.000

6 **a** 0.0003 m **b** 5×10^9 mJ
c 1×10^7 kW

2 Balancing chemical equations

1 **a** $2C + O_2 \rightarrow 2CO$ **b** $N_2 + 3H_2 \rightarrow 2NH_3$
c $C_2H_4 + 3O_2 \rightarrow 2H_2O + 2CO_2$

2 **a** $C_6H_{14} + 9\frac{1}{2}O_2 \rightarrow 6CO_2 + 7H_2O$ **or** $2C_6H_{14} + 19O_2 \rightarrow 12CO_2 + 14H_2O$
b $2NH_2CH_2COOH + 4\frac{1}{2}O_2 \rightarrow 4CO_2 + 5H_2O + N_2$
or $4NH_2CH_2COOH + 9O_2 \rightarrow 8CO_2 + 10H_2O + 2N_2$

3 **a** $Mg(OH)_2 + 2HNO_3 \rightarrow Mg(NO_3)_2 + 2H_2O$
b $3Fe(NO_3)_2 + 2Na_3PO_4 \rightarrow Fe_3(PO_4)_2 + 6NaNO_3$

3 Rearranging equations and calculating concentrations

1 **a** $n = cv$ **b** $v = \frac{n}{c}$

2 **a** $n = \frac{PV}{RT}$ **b** $T = \frac{PV}{nR}$

3 $\frac{0.2}{0.050} = 4.0$ mol dm^{-3}

4 $\frac{0.05}{2} = 0.025$ mol dm^{-3}

5 $\frac{36}{1000} \times 0.1 = 3.6 \times 10^{-3}$ mol

4 Molar calculations

1 **a** $\frac{0.486}{24.3} = 0.02$ mol **b** 0.02 mol
c $0.02 \times 40.3 = 0.806$ g

2 **a** $\frac{4.25}{85} = 0.05$ mol **b** $\frac{0.05}{2} = 0.025$ mol

3 **a** $\frac{500}{84.3} = 5.93$ mol **b** 5.93 mol

5 Percentage yields and percentage errors

1 3.19/4.75 × 100 = 67.2%

2 6.25/12.00 × 100 = 52.1%

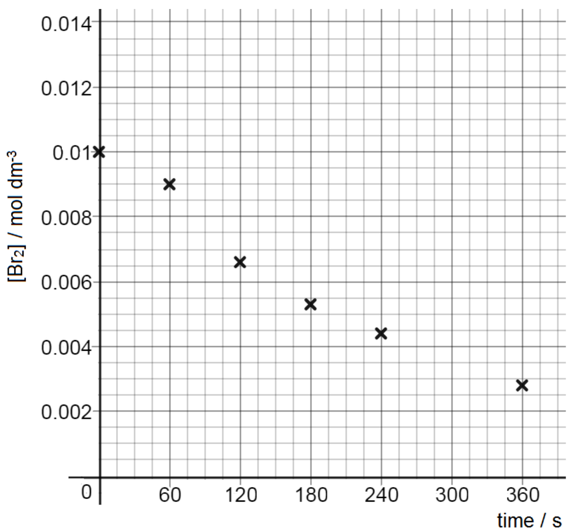
3 **a** 0.5/21 × 100 = 2.38% **b** 0.5/43 × 100 = 1.16%

4 **a** 0.5 × (2/12) × 100 = 8.33% **b** 0.5 × (2/37.6) × 100 = 2.66%

6 Graphs and tangents

1 $\frac{-1.25}{65} = -0.0192$

2 **a**



b Half-life is approximately 180 seconds **c** The reaction is first order