

**A Level Chemistry**

**Transition Work**

1. To prepare you for the first year of A’ level chemistry, you should work through ‘Head Start to AS Chemistry’ by David Mason

ISBN: 978 1 78294 2801



Available from Amazon and other booksellers

for approximately £5.

Read each section then answer questions.

Check and correct answers before moving onto the next section.

1. You also need to complete the (amended) Edexcel Transition Activity (which starts on page 11) and answer the following exam style questions. The questions are higher tier GCSE chemistry questions and are relevant to the topics covered at A’ level during the first term. **You must bring your answers to all of this along to your first chemistry lesson in September.**
2. The Text book for the course can be ordered through the College when you start in September, If you wish to purchase it before then the details are given below:

|  |  |
| --- | --- |
| Edexcel A’ Level Chemistry 1By Graham Curtis, Andrew Hunt and Graham HillISBN 978-1-4718-0746-6 | Product Details |

**Questions**

The following questions would take about 1 hour under test conditions.

However, **this is not a test, it is revision!**

* Please use your GCSE notes, ‘Head start’ or a GCSE revision guide to help.
* Take however long you need to remind yourself how to tackle these types of question!
* It may therefore take 2-3 hours to do properly.
* Please complete the table below and attempt all questions.

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| --- | --- |
| **Name** | **GCSE Grades:**  |
|  | Maths | Chemistry OR Science | Physics OR Additional Science | Biology ORN/A |
|  |  |  |  |  |

**Q1.**

Chlorine is an element in group 7 of the periodic table.

Chlorine, Cl2, is a simple molecular, covalent substance.

The atoms in a molecule of chlorine are held together by a covalent bond.

(i)  Explain what is meant by the term **covalent bond**.

**(2)**

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(ii)  Phosphorus reacts with chlorine to form phosphorus trichloride, PCl3.

A phosphorus atom has five electrons in its outer shell.

A chlorine atom has seven electrons in its outer shell.

Draw the dot and cross diagram to show the bonding in a molecule of phosphorus trichloride, PCl3.

Show outer electrons only.

**(2)**

(iii)  Aluminium reacts with chlorine to form aluminium chloride, AlCl3.

Write the balanced equation for this reaction.

**(2)**

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**(Total for question = 6 marks)**

**Q2.**An atom of copper has an atomic number of 29 and a mass number of 63.

(i)  Complete the table to show the numbers of protons, neutrons and electrons in this atom of copper.

**(2)**



(ii)  Copper is in period 4 of the periodic table.

State what information this gives about the number of shells that contain electrons, in a copper atom.

**(1)**

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(iii)  Copper exists as isotopes.

Explain what is meant by the term **isotopes**.

**(2)**

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(iv)  A sample of copper contains

70% of copper-63 atoms and

30% of copper-65 atoms.

Use this information to calculate the relative atomic mass of copper in this sample.

**(3)**

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relative atomic mass of copper = ...........................................................

 **Q3.**

(a) The table shows the number of electrons, neutrons and protons in particles P, Q, R, S, T and V.



(i) Which particle is a negatively charged ion?

Put a cross (  ) in the box next to your answer.

**(1)**

    **A**   P

    **B**   S

    **C**   T

    **D**   V

(ii) Which particles are atoms of metals?

Put a cross (  ) in the box next to your answer.

**(1)**

    **A**   P and R

    **B**   Q and R

    **C**   Q and S

    **D**   Q, S and V

(b) Each element has an atomic number.

(i) State what is meant by **atomic number.**

**(1)**

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 (ii) The atomic number of boron is 5.
       Boron exists as two isotopes boron-10 and boron-11.

Use this information to explain why boron-10 and boron-11 are isotopes.

**(2)**

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(c) (i) Explain what is meant by the term relative atomic mass.

**(2)**

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 (ii) A sample of boron contains
      19.7% of boron-10.
      80.3% of boron-11.

     Use this information to calculate the relative atomic mass of boron.

**(3)**

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 **(Total for Question = 10 marks)**

**Q4.**

The atomic number of carbon is 6.

The atomic number of hydrogen is 1.

Draw a dot and cross diagram of a molecule of methane, CH4.

Show the outer shell electrons only.

**(2)**

**Q5.**

Marble chips react with dilute hydrochloric acid.

Marble is a form of calcium carbonate.

(i)  Complete the balanced equation for this reaction.

**(2)**



(ii)  Explain how using smaller sized marble chips affects the rate of this reaction, when all the other conditions remain the same.

**(2)**

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(iii)  Explain, in terms of collisions between particles, how increasing the concentration of the hydrochloric acid affects the rate of this reaction, when all the other conditions remain the same.

**(2)**

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**(Total for question = 6 marks)**

**Q6.**

(i)  When calcium carbonate is heated, it breaks down to form calcium oxide and carbon dioxide.

What type of reaction is this?

Put a cross () in the box next to your answer.

**(1)**

   **A**    combustion

   **B**    decomposition

   **C**    oxidation

   **D**    precipitation

(ii)  Calcium oxide reacts with water to form calcium hydroxide, Ca(OH)2.

Write the balanced equation for the reaction between calcium oxide and water.

**(2)**

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**(Total for question = 3 marks)**

**Q7.**

In the extraction of titanium from its ore, the final stage involves the reaction between titanium(IV) chloride, TiCl4, and sodium.

TiCl4 + 4Na → Ti + 4NaCl

Calculate the maximum mass of titanium that can be obtained from 500 tonnes of titanium(IV) chloride in this reaction.

(relative atomic mass: Ti = 48
relative formula mass of TiCl4 = 190)

**(2)**

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 mass of titanium = ........................................................... tonnes

**(Total for question = 2 marks)**

**Q8.**

Calculate the percentage by mass of nitrogen in ammonium nitrate, NH4NO3.

(relative atomic masses: H = 1.0, N = 14, O = 16)

**(3)**

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 percentage by mass of nitrogen = ........................................................... %

**(Total for question = 3 marks)**

**Q9.**

When nitrogen and hydrogen react to form ammonia, the reaction can reach a dynamic equilibrium.



    (i) Calculate the minimum volume of hydrogen required to completely convert 1000 dm3 of nitrogen into ammonia.

**(1)**

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volume of hydrogen =. . . . . . . . . . . . . . . . . .dm3

 (ii) Ammonia is reacted with excess nitric acid, HNO3, to make ammonium nitrate, NH4NO3.

NH3 + HNO3 → NH4NO3

Calculate the mass of ammonium nitrate produced by the complete reaction of 34 g of ammonia.

(Relative atomic masses H = 1.0, N = 14, O = 16)

**(3)**

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mass of ammonium nitrate produced =. . . . . . . . . . . . . . . . .g

**Q10.**

In an experiment, 3.1 g of phosphorus reacted with 24 g of bromine to form phosphorus bromide.

Calculate the empirical formula of the phosphorus bromide.

You must show your working.

(relative atomic masses: P = 31, Br = 80)

**(3)**

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empirical formula ...........................................................

**Q11.**

Copper nitrate contains copper ions, Cu2+, and nitrate ions, .

(i)  Describe, in terms of electrons, how a copper atom, Cu, becomes a copper ion, Cu2+.

**(2)**

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(ii)  Write the formula for copper nitrate.

**(1)**

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**Q12.**

Aluminium ions, Al3+, react with hydroxide ions in solution to give a white precipitate of aluminium hydroxide.

Write the ionic equation for this reaction.

**(3)**

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**Q13.**

(a)  The halogens react with hydrogen to form hydrogen halides.

Complete the balanced equation for the reaction between hydrogen and bromine forming hydrogen bromide.

**(2)**

H2 + Br2 → ...........................................................

(b)  Calculate the relative formula mass of magnesium chloride, MgCl2.
(relative atomic masses: Mg = 24.0; Cl = 35.5)

**(1)**

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relative formula mass = ...........................................................

(c)  Calculate the percentage by mass of fluorine in sodium fluoride, NaF.
(relative atomic masses: F = 19; Na = 23)

**(2)**

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percentage by mass of fluorine = ...........................................................

EDEXCEL A LEVEL Transition Activity

Introduction

Student’s Checklist for Being A Level Chemistry-ready

You are expected to know/understand the following:

**Electron configuration** for the first 20 elements

**Naming compounds** from formulae and vice versa

**Bonding**

Three main types – formation and properties

* ionic
* covalent
* and metallic

**Dot and cross diagrams** for covalent molecules and ionic compounds to include:

* sodium chloride, calcium oxide, calcium fluoride, aluminium oxide
* chlorine, oxygen, nitrogen, ammonia, carbon dioxide, methane, ethane, ethanol, sulfur dioxide, water

**Writing formulae** for all of the above plus compounds with:

* carbonate
* nitrate
* hydroxide
* hydrogen carbonate
* sulfate.

**Balancing equations** for

* neutralisation
* metals with acids
* alkali metals with water
* redox (displacement of halogens and metals), thermal decomposition

**Calculations**

* relative atomic mass, relative formula mass and empirical formulae
* Percentage yield and atom economy
* Reacting masses and limiting reagent

**Energetics**

* difference between exothermic and endothermic
* graphs associated with these
* energies in bond making and bond breaking

**Organic Chemistry**

* differences between alkanes and alkenes
* naming and reactions of alkanes and alkenes
* fractional distillation
* cracking
* characteristics of good fuels
* balancing combustion equations

Section A: Atomic structure, formulae and bonding

This section reviews the fundamental concepts from GCSE Topic 1: Key Concepts. It is important to emphasise that the AS concepts are amplifications of what was learned at KS4.

Students’ strengths and common misconceptions

The table below outlines the areas in which most students do well and the common mistakes and misconceptions across the topics listed.

|  |  |  |
| --- | --- | --- |
|  | Strengths | Common mistakes |
| Atomic structure | Listing subatomic particles and their properties (mass and charge). | Being unclear about Subatomic particles in ions. |
| Electron configuration | Simple 2.8.8 rule. | Deducing group number for the p-block elements.Misunderstanding electron configuration for ions. |
| Dot-and-cross diagrams | Knowing the general rule for individual atoms.Simple ionic compounds, e.g. NaCl. | Checking the total outer electrons after bonding – both ionic and covalent.Overlapping shells for ionic compounds.Missing charges on ions. |

Summary sheets

KS4 – Atomic structure

**Subatomic particles:** nucleus (protons and neutrons), electrons in shells.

Describe the particles in terms of their relative masses and relative charges:

* Protons – mass 1, charge +1.
* Electrons – mass = negligible ( ), charge –1.
* Neutrons – mass = 1, charge = 0.

**Notes**

* Number of protons = number of electrons (uncharged/neutral atoms).
* Proton number = atomic number.
* Mass number = protons + neutrons.

KS4 – Isotopes and calculating relative isotopic mass

Isotopes are *atoms* of the same elements which have different numbers of *neutrons* but the same number of *protons*.



Relative isotopic mass =

KS4 – Ionic compounds

Formation of ions

Atoms of metallic elements in Groups 1,2 and 3 can form positive ions when they take part in reactions since they are readily able to lose electrons.

Atoms of Group 1 metals lose one electron and form ions with a 1+ charge, e.g. Na+

Atoms of Group 2 metals lose two electrons and form ions with a 2+ charge, e.g. Mg2+

Atoms of Group 3 metals lose three electrons and form ions with a 3+ charge, e.g. Al3+

Atoms of non-metallic elements in Groups 5, 6 and 7 can form negative ions when they take part in reactions since they are able to gain electrons.

Atoms of Group 5 non-metals gain three electrons and form ions with a 3– charge, e.g. N3–

Atoms of Group 6 non-metals gain two electrons and form ions with a 2– charge, e.g. O2–

Atoms of Group 7 non-metals gain one electrons and form ions with a 1– charge, e.g. Cl–

![C:\Users\Samia.ElAli\AppData\Local\Microsoft\Windows\Temporary Internet Files\Content.Word\images[2].jpg]()

ANions = Negative Ca+ions = +ive

Why are ions negative or positive?

* Find the atomic number (the smaller number with the symbol).
* This equals the number of protons, which equals the number of electrons in an uncharged/neutral atom.
* If electrons are lost from the atom, there are now more protons than electrons, so the ion is positively charged.
* If electrons are gained by the atom, there are now fewer protons than electrons, so the ion is negatively charged.

KS4 – Electron configuration

Filling electron shells

* *n* = 1, maximum = 2e–
* *n* = 2; maximum = 8e–
* *n* = 3 ;maximum = 18e–
* *n* = 4; maximum = 32e–

**Representing electron configurations**

* Write as, e.g. 2.8.3 or 2,8,3

Using the Periodic Table

* Period number (row) = number of shells
* Group number (column) = number of electrons in the outer (last) shell

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| Group number | 1 | 2 | 3 |  | 5 | 6 | 7 |
|  | **Li** |  | **Be** |  | **B** |  |  |  | **N** |  | **O** |  | **F** |  |
|  | Atom | Ion | Atom | Ion | Atom | Ion |  |  | Atom | Ion | Atom | Ion | Atom | Ion |
| **Electrons** | –3 | –2 | –4 | –2 | –5 | –2 |  |  | –7 | –10 | –8 | –10 | –9 | –10 |
| **Protons** | +3 | +3 | +4 | +4 | +5 | +5 |  |  | +7 | +7 | +8 | +8 | +9 | +9 |
| **Overall charge** | 0 | 1+ | 0 | 2+ | 0 | 3+ |  |  | 0 | 3– | 0 | 2– | 0 | 1– |
| **Electron configuration** | 2.1 | 2 | 2.2 | 2 | 2.3 | 2 |  |  | 2.5 | 2.8 | 2.6 | 2.8 | 2.7 | 2.8 |
| **Name of ions** | lithium | beryllium | boron |  | nitride | oxide | fluoride |
|  | Lose electrons, charge = +group number | Gain electrons, charge = group number – 8  |

KS4 – Dot-and-cross diagrams for ionic bonding

Hints and tips

Always …

… count the electrons!

… remember that ions should have full outer shells.

… make sure that when an ion is formed, you put square brackets round the diagram and show the charge.

Never …

… show the electron shells overlapping.

… show electrons being shared (ions are formed by the **transfer** of electrons!).

… remove electrons from the inner shell.

… give metals a negative charge.





KS4 – Covalent compounds (simple covalent bonding)

Hints and tips

Always …

… show the shells touching or overlapping where the covalent bond is formed.

… count the final number of electrons around each atom to make sure that the outer shell is full.

Never …

… include a charge on the atoms.

… draw the electron shells separated.

… draw unpaired electrons in the region of overlap.

The two diagrams below only show the outer-shell electrons.



Worksheet 1: Atomic structure and the Periodic Table

Complete the following sentences and definitions to give a summary of this topic.

Structure of an atom

The nucleus contains …

The electrons are found in the …

To work out the number of each sub-atomic particle in an atom we use the Periodic Table (PT). The number of protons is given by …

In a neutral atom the number of electrons is …

To work out the number of neutrons we …

Vocabulary

**State what is meant by the following terms.**

1. Relative atomic mass
2. Relative molecular mass
3. Isotope
4. Relative isotopic mass

Structure of an ion

When an atom becomes an ion, only the number of changes.

For positive ions this by the number equivalent to the charge on the ion.

For negative ions this by the number equivalent to the charge on the ion.

Exam practice

GCSE/GCE Overlap

1. The relative atomic mass of an element is determined using a mass spectrometer.

State what is meant by the term *relative atomic mass*.

(2 marks)

(Edexcel GCE Jan 2011, 6CH01, Q15a)

1. Chlorine forms compounds with magnesium and with carbon.
2. Draw a dot-and-cross diagram to show the electronic structure of the compound magnesium chloride (only the outer electrons need be shown). Include the charges present.

(2 marks)

1. Draw a dot-and-cross diagram to show the electronic structure of the compound tetrachloromethane (only the outer electrons need be shown).

(2 marks)

1. Draw a dot-and-cross diagram of a molecule of carbon dioxide. Show outer electrons only.

(2 marks)

(Edexcel GCSE Jun 2013, 5CH2H, Q2(iii))

**3** **a** State what is meant by the term *relative isotopic mass*.

(2 marks)

**b** State what is meant by the term *isotopes*.

(2 marks)

**c i** State what is meant by the term *relative atomic mass*.

(2 marks)

**ii** A sample of boron contains:

* 19.7% of boron-10
* 80.3% of boron-11.

Use this information to calculate the relative atomic mass of boron.

(3 marks)

(Edexcel GCSE May 2013, 5CH2H, Q4c(i)–(ii))

1. A molecule is …

**A** a group of atoms joined by ionic bonding.

**B** a group of atoms joined by covalent bonding.

**C** a group of ions joined by covalent bonding.

**D** a group of atoms joined by metallic bonding.

(1 mark)

1. The relative atomic mass is defined as …

****

**A** the mass of an atom of an element relative to the mass of a carbon-12 atom.

**B** the mass of an atom of an element relative to the mass of a hydrogen atom.

**C** the average mass of an element relative to the mass of a carbon atom.

**D** the average mass of an atom of an element relative to the mass of a
carbon-12 atom.

(1 mark)

(Edexcel GCE Jan 2012, 6CH01, Q1,2)

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1. The isotopes of magnesium Mg and Mg both form ions with charge 2+. Which of the following statements about these ions is true?

**A** Both ions have electronic configuration 1s2 2s2 2p6 3s2.

**B**  Mg2+ has more protons than Mg2+.

**C** The ions have the same number of electrons but different numbers of neutrons.

**D** The ions have the same number of neutrons but different numbers of protons.

(1 mark)

1. ****The radioactive isotope iodine-131, I, is formed in nuclear reactors providing nuclear power. Naturally occurring iodine contains only the isotope I.

**a** Complete the table to show the number of protons and neutrons in these two isotopes.



|  |  |  |
| --- | --- | --- |
| Isotope |  I |  I |
| Number of protons |  |  |
| Number of neutrons |  |  |

(2 marks)

**b** When iodine-131 decays, one of its neutrons emits an electron and forms a proton. Identify the new element formed.

(1 mark)

(Edexcel GCE May 2013, 6CH01R, 18a,b)

Section B: Quantitative analysis and equations

This section covers one of the most important areas of the chemistry specification. A good understanding of the concepts covered here, particularly reacting masses, will have a huge impact on your study of later topics, including the A2 specification. The Table below lists the areas that students most commonly struggle with.

Unlike Physics, formulae and equations are not provided in Chemistry exams so it is important that students know these very well and, more importantly, are able to manipulate them as necessary to solve a given problem. In addition to this, students will need to decide on the appropriate number of significant figures to use in their final answer.

Students’ strengths and common misconceptions

The table below outlines the areas in which most students do well and the common mistakes and misconceptions across the topics listed.

|  |  |  |
| --- | --- | --- |
|  | Strengths | Common mistakes |
| Quantities in chemistry | Definitions as ‘standalone’. | Conversions from one quantity to another, e.g. moles to grams.Not recognising that molar quantities are the same but the method of calculation depends on the species, e.g. solutes in solution, gases, solids. |
| Empirical formulae | Writing empirical formula from molecular formula.Recognising a mathematical relationship between % composition and *A*r. | Inverting the %/*A*r ratio.Failing to simplify the ratios.Writing a final answer.Deducing molecular formula from empirical formula and *M*r. |
| Balancing equations | Simple acid–alkali and metal plus oxygen or halogen equations. | Translating practical scenarios into word and formula equations.Not learning the common ‘known’ reactions, e.g. carbonate plus acid.Applying the law of conservation of mass to equations. |
| Ionic equations | Given the state symbols, be able to split the ions. | Not knowing which species are soluble and the state symbols of common chemical species.Splitting common acids. |
| Reacting masses | Conservation of mass.Working out masses or moles as standalone direct questions. | Selecting the correct formula when solving problems with practical scenarios.Following multistep procedures and calculations. |

Summary sheet: Writing formulae

Writing formulae

Compounds should have no overall charges, so the positive and negative charges should cancel each other out.

Apart from working out the charges on ions made up of one element, you need to know the following compound ions and their charges.

|  |  |  |
| --- | --- | --- |
| Name | Formula | Charge |
| hydroxide | OH– | 1– |
| nitrate | NO3– | 1– |
| sulfate | SO42– | 2– |
| carbonate | CO32– | 2– |
| ammonium | NH4+ | 1+ |

Follow these steps.

|  |  |  |
| --- | --- | --- |
| Write the name of the compound | Magnesium bromide | Sodium sulfate |
| Work out the charge of your positive ion = group number, or 1+ for ammonium. | Mg2+ | Na+ |
| Work out the charge of your negative ion = group number – 8 *or* known charge for a compound ion. | Br– | SO42– |
| Rewrite the symbols; put a bracket around any compound ion. | Mg2+ Br–Mg Br | Na+ SO42–Na (SO4) |
| Swap the numbers of the charges and drop them to the opposite ion. | MgBr2 | Na2(SO4) |

Writing ionic equations

* Make sure all state symbols are included.
* Identify the species that are aqueous, using the rules of solubility.
1. Look at the cation – is it Group 1 or ammonium? If so → soluble.
2. Look at the anion – is it a nitrate? If so → soluble.
* Proceed only if you have ruled out 1 and 2.
1. Is the anion a halide (chloride, bromide or iodide)?
2. If so, look at the metal – lead or silver? If so → insoluble.
3. Is the anion a sulfate?
4. If so, look at the metal – barium, calcium, lead? If so → insoluble.
5. Is the anion a hydroxide?
6. If so, look at the metal – transition metal or Group 2 (after Ca)?
 If so → insoluble.
* Split all the soluble salts into their aqueous ions on both sides – remember to write the numbers in front of the ions for multiples.
* Cancel out the ions that appear on both sides – again pay attention to numbers.
* Write your final equation (always keep the state symbols unless specifically told
not to!).

Reacting masses

To work out masses of reactants and products from equations, follow these steps.

|  |  |  |
| --- | --- | --- |
| Steps to follow | Example | Example |
|  | 5 g of Ca reacted with excess chlorine. What mass of CaCl2 is formed? | When MgCO3 was heated strongly, 4 g of MgO was formed. What is the mass of MgCO3 that was heated? |
| Write the balanced equation. | Ca + Cl2 → CaCl2 | MgCO3 → MgO + CO2(g) |
| Write the masses given. | 5 g (excess)  **?** |  **?** 4 g |
| Find the *A*r or *M*r. | 40 111 |  84 40 |
| Divide by the atomic or molecular mass (step 2 ÷ step 3). |  :  |  :  |
| Treat these like ratios, rearrange to find the unknown (**?**). | Mass of CaCl2 = (5 × 111) ÷ 40 = 13.9 g | Mass of MgCO3 = (4 × 84) ÷ 40 = 8.4 g |

Note: if you are told something is in excess, do not use it in the calculation!

Percentage yield

The calculations above dealt with the masses you get or use if the reaction is 100% complete.

Most reactions are not 100% complete for the following reasons:

* not all the reactant reacts
* some is lost in the glassware as you transfer the reactants and the products
* some other products might be formed that you do not want.

This is a problem in industry. Less of the desired product has been made, so there is less to use or sell, and the waste has to be disposed of. Waste products can be harmful to the environment, e.g. the one above produces the greenhouse gas CO2. Industries try to choose reactions that minimise waste and do not produce harmful products. They also try to make the rate of reaction high enough to make the reaction turnover fast so they can increase production and make money.

To work out % yield: use the balanced equation to work out how much of the given product you should get if the reaction is 100% efficient – this is the theoretical yield.

Then: 

Worksheet 1: Chemical formulae

**Write the formulae of the following compounds.**

|  |  |
| --- | --- |
| Copper(II) sulfate |   |
| Nitric acid |   |
| Copper(II) nitrate |   |
| Sulfuric acid |   |
| Sodium carbonate |   |
| Aluminium sulfate |   |
| Ammonium nitrate |   |
| Nitrogen dioxide |   |
| Sulfur dioxide |   |
| Ammonia |   |
| Ammonium sulfate |   |
| Potassium hydroxide |   |
| Calcium hydroxide |   |

Worksheet 2: Cations and anions

**Complete the table below to show the substance, its formula and its
individual ions.**

|  |  |  |  |
| --- | --- | --- | --- |
| Substance | Formula | Cation (exact number) | Anion (exact number) |
| Sodium bromide |  |  |  |
|  | KI |  |  |
| Silver nitrate |  |  |  |
| Copper(II) sulfate |  |  |  |
|  | NaHCO3 |  |  |
| Magnesium carbonate |  |  |  |
| Lithium carbonate |  |  |  |
|  | Ca(HSO4)2 |  |  |
| Aluminium nitrate |  |  |  |
| Calcium phosphate |  |  |  |
| Potassium hydride |  |  |  |
| Sodium ethanoate |  |  |  |
|  | KMnO4 |  |  |
| Potassium dichromate(VI) |  |  |  |
| Zinc chloride |  |  |  |
| Strontium nitrate |  |  |  |
| Sodium chromate(VI) |  |  |  |
| Calcium fluoride |  |  |  |
| Potassium sulfide |  |  |  |
| Magnesium nitride |  |  |  |
| Lithium hydrogensulfate |  |  |  |
|  | (NH4)2SO4 |  |  |

Worksheet 3: Writing equations

**Write: (a) the chemical equation and (b) the ionic equation for each of the following reactions.**

1. Magnesium with sulfuric acid
2. Calcium carbonate with nitric acid
3. Hydrochloric acid with sodium hydroxide
4. Aqueous barium chloride with aqueous sodium sulfate
5. Aqueous sodium hydroxide with sulfuric acid
6. Aqueous silver nitrate with aqueous magnesium chloride
7. Solid magnesium oxide with nitric acid

**8** Aqueous copper(II) sulfate with aqueous sodium hydroxide

**9** Aqueous lead(II) nitrate with aqueous potassium iodide

**10** Aqueous iron(III) nitrate with aqueous sodium hydroxide

Exam practice

1. Coral reefs are produced by living organisms and predominantly made up of calcium carbonate. It has been suggested that coral reefs will be damaged by global warming because of the increased acidity of the oceans due to higher concentrations of carbon dioxide.

**a** Write a chemical equation to show how the presence of carbon dioxide in water results in the formation of carbonic acid. State symbols are not required.

(1 mark)

**b** Write the ionic equation to show how acids react with carbonates. State symbols are not required.

(2 marks)

1. One method of determining the proportion of calcium carbonate in a coral is to dissolve a known mass of the coral in excess acid and measure the volume of carbon dioxide formed.

In such an experiment, 1.13 g of coral was dissolved in 25 cm3 of hydrochloric acid (an excess) in a conical flask. When the reaction was complete, 224 cm3 of carbon dioxide had been collected over water using a 250 cm3 measuring cylinder.

**a** Draw a labelled diagram of the apparatus that could be used to carry out this experiment.

(2 marks)

**b** Suggest how you would mix the acid and the coral to ensure that no carbon dioxide escaped from the apparatus.

(1 mark)

**c** Complete the equation below for the reaction between calcium carbonate and hydrochloric acid by inserting the missing state symbols.

CaCO3(........) + 2HCl(........) → CaCl2(........) + H2O(l) + CO2(........)

(1 mark)

**d** Calculate the molar mass of calcium carbonate. (Assume relative atomic masses: Ca = 40.1, C = 12.0, O = 16.0)

(1 mark)

1. Magnesium chloride can be made by reacting solid magnesium carbonate, MgCO3, with dilute hydrochloric acid.

**a** Write an equation for the reaction, including state symbols.

(2 marks)

**b** A precipitate of barium sulfate is produced when aqueous sodium sulfate is added to aqueous barium chloride. Give the ionic equation for the reaction, including state symbols.

(2 marks)

Section C: Structure and properties – Literacy Focus

In this section we apply the concepts covered in Section A to properties of materials.

Students’ strengths and common misconceptions

The table below outlines the general areas in which students do well and the common mistakes and misconceptions across the topics listed.

|  | What most students can do (well) | Common mistakes |
| --- | --- | --- |
| **Metals** | Stating the physical properties of metals, including conductivity.Describing the structure as particles with delocalised electrons. | Using words like *molecules* and *atoms* instead of *cations* or *ions*, and *free* instead of *delocalised* or *free-moving* electrons to describe metallic bonds.Explaining the differences in the melting point and electrical conductivity of two metals.Describing metallic structure as ‘protons’ in a sea of electrons |
| **Ionic compounds** | Knowing that ionic compounds form giant structures, and therefore have high melting points.Knowing that ionic compounds conduct electricity when molten or in solutions. | In explaining or describing the electrostatic attraction between cations and anions in the giant structure.When describing separation of the ions at melting temperature.Explaining why ionic compounds conduct electricity when molten or in solution using terms like *free electrons* instead of in terms of mobility of ions. |
| **Covalent compounds** | Knowing the existence of simple molecular and giant covalent structures and give examples of each.Knowing of the existence of intermolecular forces and the effect of increasing molecular mass.In diamond each carbon atom forms covalent bonds with four others whereas in graphite it bonds only with three. | Explaining the boiling point – distinguishing between intermolecular forces in simple molecules and extensive covalent bonds in giant structures.Use of terms like ‘carbon bonds’ instead of covalent bonds |

Summary sheet 1: Structure and bonding

Words used to describe structure and bonding:

* ions, atoms, molecules, intermolecular forces, electrostatic forces, delocalised electrons, cations, anions, outer electrons, shielding

Metallic bond: electrostatic attraction between the nuclei of cations (positive ions) and delocalised electrons.

Strength of the metallic bonding increases with the number of valence electrons (outer electrons in the atoms) and with decreasing size of the cation.

Ionic bonds and ionic compounds

Explain why NaCl has a high melting point and only conducts electricity when molten or in solution. (6 marks)

**An answer should cover the following points.**

1. The Na+ and Cl– ions are held by strong electrostatic forces.
2. To melt solid NaCl, energy is needed to separate overcome the forces of attraction sufficiently for the lattice structure to break down and for the ions to be free to slide past one another.
3. Even though the ions are charged, the solid cannot conduct electricity because the ions are not mobile (free to move).
4. If the solid is melted, the ions can move freely and allow the liquid to conduct electricity.
5. Also, when dissolved in water the *ions* are separated by the water molecules and so are free to move, hence the aqueous solution can conduct electricity.

 

Summary sheet 2: Diamond and graphite structures



|  |  |  |
| --- | --- | --- |
| Property | Diamond | Graphite |
| Melting point | High – atoms held by strong covalent bonds.Many covalent bonds must be broken to melt it.Lots of energy required.Is solid at room temp. | High – atoms held by strong covalent bonds.Many covalent bonds must be broken to melt it.Lots of energy required.Is a solid at room temp. |
| Electrical conductivity | Poor – no mobile electrons available.All four outer electrons of each carbon are used in bonding. | Good – each carbon only uses three of its outer electron to form covalent bonds. 4th electron form each atom contributes to a delocalised electron system. These delocalised electrons can flow when a potential difference is applied parallel to the layers. |
| Lubricant | Poor – structure is rigid. | Gas molecules are trapped between the layers and allow the layers to slide past one another.Same reason for its use in pencils. |
| Solubility | Insoluble in water – no charged particles to interact with water (think of SiO2, main component of sand). | Insoluble in water – no charged particles to interact with water (think of SiO2, main component of sand). |

Using key words to describe ionic structure

Describe and explain how the structure of sodium fluoride is formed.

**Use knowledge of the structure of sodium chloride**



**Which key words will you need?**

* Attraction
* Electrostatic
* Tight
* Non-metals
* Giant
* Packed
* Anions
* Strong
* Metals
* Forces
* Ionic
* Opposition
* Lattice
* Cations

**Tip**

For questions about the physical properties of ionic compounds, relate the properties to their bonding and structure.

|  |  |
| --- | --- |
| Property | Why? |
| Does not conduct electricity when solid. |  |
| Conducts electricity when molten or in aqueous solution. |  |
|  | The ions are held by strong electrostatic forces of attraction and a large amount of energy is needed to overcome the attractions. |
|  | The ions are tightly packed together. |

Exam practice

1. Suggest why the melting temperature of magnesium oxide is higher than that of magnesium chloride, even though both are almost 100% ionic.

(3 marks)

**Edexcel GCE Jan 2011, 6CH01, Q17**

1. Silicon exists as a giant covalent lattice.

**a** The electrical conductivity of pure silicon is very low. Explain why this is so in terms of the bonding.

(2 marks)

**b** Explain the high melting temperature of silicon in terms of the bonding.

(2 marks)

**Edexcel GCE Jan 2012, 6CH01**

1. The melting temperatures of the elements of Period 3 are given in the table below. Use these values to answer the questions that follow.



**a** Explain why the melting temperature of sodium is very much less than that of magnesium.

(3 marks)

**b** Explain why the melting temperature of silicon is very much greater than that of white phosphorus.

(3 marks)

**c** Explain why the melting temperature of argon is the lowest of all the elements of Period 3.

(1 mark)

**d** Explain why magnesium is a good conductor of electricity whereas sulfur is a non-conductor.

(2 marks)

Baseline assessment questions

1. Give the formulae of the following compounds.

|  |  |
| --- | --- |
| Copper(II) sulfate Sodium hydroxide Strontium nitrate Sodium carbonate  | Lithium hydrogencarbonate Potassium nitrate Calcium hydroxide Aluminium fluoride  |

(4 marks)

1. Name the following compounds.

|  |  |
| --- | --- |
| NH4Cl C2H4  CO2  Fe2O3  HBr  | HNO3 C3H8 C2H5OH SO2 NH3  |

(5 marks)

1. Complete the table below.

|  |  |  |  |
| --- | --- | --- | --- |
| Particle | Where it is found | Charge | Mass |
|  |  | 0 |  |
| Proton |  |  |  |
|  |  |  | 0 |

(3 marks)

1. Deduce the relative formula mass of the following.

|  |  |
| --- | --- |
| SO2 C2H6 C2H5OH NH4Cl  | KBr Ca(OH)2 NaNO3 FeCl3  |

(4 marks)

1. State what is meant by the following terms.
2. the mass number of an atom

(1 mark)

1. relative atomic mass

(2 marks)

1. isotopes

(2 marks)

1. For the following reactions, write:
2. the word equation (1 mark)
3. the chemical equation complete with state symbols. (2 marks)

|  |
| --- |
| Calcium carbonate and hydrochloric acid |
| Magnesium and sulfuric acid |
| Complete combustion of butane |
| Thermal decomposition of calcium carbonate |
| Sodium and water |

(12 marks)

1. State what is meant by the following terms.

Ionic bonding

Covalent bonding

Metallic bonding

(3 marks)

1. Complete the table below. You may use the following words to help you.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| ionic | covalent | giant | simple | metallic |

|  |  |  |  |
| --- | --- | --- | --- |
| Substance | Formula | Type of bonding | Type of structure |
| Hydrogen sulfide |  |  |  |
| Graphite |  |  |  |
| Silicon dioxide |  |  |  |
| Methane |  |  |  |
| Calcium |  |  |  |
| Magnesium chloride |  |  |  |

(6 marks)

1. Explain why graphite can be used as a solid lubricant and also as electrodes.
2. marks)

13 Complete the table below.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | Number of protons | Number of electrons | Number of neutrons | Electron configuration |
| S |  |  |  |  |
| Mg |  |  |  |  |
| O2– |  |  |  |  |
| H+ |  |  |  |  |
| Kr |  |  |  |  |
| Al3+ |  |  |  |  |

1. marks)

14 Draw a dot-and-cross diagrams for the following compounds.

**a** Methane

**b** Water

**c** Sodium fluoride

**d** Magnesium bromide

**e** Ammonia

**f** Potassium oxide

**g** Calcium oxide

**h** Oxygen

**i** Carbon dioxide

(18 marks)

1. solid potassium carbonate was reacted with excess hydrochloric acid, and the change in mass was recorded as shown in the diagram below.



The equation for the reaction is given by:

K2CO3(s) + 2HCl(aq) 🡪 2KCl(aq) + H2O(l) + CO2(g)

**a** Write the ionic equation for this reaction.

**b** Calculate the relative molecular mass *M*r of K2CO3.

(2 marks)

1. Complete the Table below using the following words:

ionic covalent giant simple metallic

|  |  |  |  |
| --- | --- | --- | --- |
| Substance | Formula | Type of bonding | Structure |
| Hydrogen sulfide |  |  |  |
| Graphite |  |  |  |
| Silicon dioxide |  |  |  |
| Calcium |  |  |  |
| Magnesium chloride |  |  |  |
| Fluorine |  |  |  |
| Argon |  |  |  |

(7 marks)

1. By considering the type of bonding and structure, explain why aluminium melts at a higher temperature than lithium.

(3 marks)

1. Explain why potassium chloride does not conduct electricity when solid whereas copper does