***EXAM DATE:***

***Teacher: \_\_\_\_\_\_\_\_\_\_\_\_\_\_***

***Name: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_***

**C3 – Quantitative Chemistry**

Foundation pupils will need to look carefully through the entire booklet. Most of C3 is for higher or triple pupils only, so care must be taken not to read anything irrelevant. It may be useful to cross anything you don’t need out!

**The conservation of mass**

In a chemical reaction, the total mass of the product is equal to the total mass of the reactants. This is called the conservation of mass.

Mass is conserved (kept the same) because no atoms are lost or made. Chemical symbol equations must always be balanced to show this, i.e. there must be the same number of atoms of each element of both sides of the equation.

For example:

Fe(s) + CuSO4(aq) 🡪 Cu(s) + FeSO4(aq)

As you can see, there is 1 atom of iron (Fe) on each side, 1 atom of copper (Cu) on each side, 1 atom of sulfur (S) on each side and 4 oxygens (O) on each side. The mass on each side is the same, therefore the mass has been conserved.

**Half equations**

This is for **higher tier ONLY**

A half equation can be used to show what happens to one reactant in a chemical reaction, with electrons written as e-.

The balanced symbol equation for the reaction between iron and copper (II) sulfate can be split into two half equations;

* The iron atom loses two electrons to form Fe2+ ions

Fe(s) 🡪 Fe2+(aq) + 2e-

* The Cu2+ ions gain two electrons to form copper atoms

Cu2+(aq) + 2e- 🡪 Cu(s)

**Ionic equations**

This is for **higher tier ONLY**

Ionic equations can be used to simplify complicated equations.

They just show the species that are involved in the reaction. In chemistry, the term species refers to the different atoms, molecules or ions that are involved in a reaction.

The spectator ions (ions not involved in the reaction) are not included. They remain unchanged in the reaction.

For example, when silver nitrate solution is added to sodium chloride solution, a white precipitate of silver chloride is produced:

AgNO3(aq) + NaCl(aq) 🡪 AgCl(s) + NaNO3(aq)

In this reaction, the nitrate ions and the sodium ions are spectator ions, so the ionic equation is:

AgNO3(aq) + NaCl(aq) 🡪 AgCl(s) + NaNO3(aq)

Ag+(aq) + Cl-(aq) 🡪 AgCl(s)

**Relative formula mass**

The relative formula mass (Mr) of a compound is the sum of the relative atomic masses (Ar) of all the atoms in the numbers shown in the formula. It is the total mass of the Ar in a compound. It doesn’t have a unit.

The Ar of the atoms are shown in the periodic table.

The Mr of calcium nitrate, Ca(NO3)2 is:

Ar Ca x 1 = 40 x 1 = 40

Ar N x 2 = 14 x 2 = 28

Ar O x 6 = 16 x 6 = 96

Mr Ca(NO3)2 = 40 + 28 + 96 = 164

The Mr of carbon dioxide (CO­2) is the sum of the Ar of 1 carbon atom and 2 oxygen atoms.

Ar C x 1 = 12 x 1 = 12

Ar O x 2 = 16 x 2 = 32

Mr CO2 = 12 + 32 = 44

**Apparent changes in mass**

Some reactions appear to involve a change in mass. This happens when reactions are carried out in a non-closed system and include a gas that can enter or leave. For example, when magnesium is burned in air

to produce magnesium oxide, the mass of the

solid increases. This is because when the

magnesium is burned it combines with oxygen

from the air and the oxygen has a mass:

2Mg(s) + O2(g) 🡪 2MgO(s)

If the mass of oxygen is included, the total mass of all the reactants is equal to the total mass of all the products.

When calcium carbonate is heated, it decomposes to form calcium oxide and carbon dioxide. The mass of the solid decreases because one of the products is a gas, which escapes into the air.

CaCO3(s) 🡪 CaO(s) + CO2(g)

If the mass of carbon is included, the total mass of all the reactants is equal to the total mass of the products.



**Quick test!**

1. State the law of conservation of mass.

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1. Why must a symbol equation be balanced?

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1. Calculate the Mr of calcium carbonate, CaCO3

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1. Why does the mass of iron oxide increase when iron is burned in air?

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**Amount of substance**

This is for **higher tier ONLY**

A **mole** (mol) is a measure of the number of particles (atoms, ions or molecules) contained in a substance.

One mole of any substance (element or compound) contains the same number of particles: six hundred thousand billion billion, or 6.02x1023. This value is known as the Avogadro constant.

The mass of one mole of a substance is its relative atomic mass or relative formula mass, in grams. For example:

One mole of sodium atoms contains 6.02x1023 atoms.

The relative atomic mass of sodium is 23.0.

One mole of sodium atoms has a mass of 23.0g.

**Calculating the amount of substance**

This is for **higher tier ONLY**

You can calculate the number of moles in a given mass of a substance using the following formula:

Amount (mol) = mass of a substance (g)

Atomic (or formula) mass (g/mol)

For example, in exams you may be given the name of a substance, and the mass of the substance. A common example would be “Calculate the number of moles of carbon dioxide in 33g of the compound”.

Mass of CO2 = 33g

Relative formula mass (Mr) of CO2 = 12 + (16 x 2) = 44

33 = 0.75 mol

44 ………………

**Balanced equations**

This is for **higher tier ONLY**

The numbers needed to balance an equation can be calculated from the masses of the reactants and the products using moles.

Aluminium oxide can be reduced to produce aluminium:

2Al2O3 🡪 4Al + 3O2

Another common question could be “Calculate the mass of aluminium oxide needed to produce 540g of aluminium”.

First, we need to use the equation to see how much aluminium oxide is needed to produce aluminium. We can see that 2 moles of Al2O3 is needed to produce 4 moles of Al. Simplified, this shows that **1 mole of Al2O3 is needed to produce 2 mole of Al.**

Second, we need to calculate the amount of aluminium there is.

Amount of aluminium = Mass

Ar of Aluminium

540 = **20mol**

27 ……………

Third, we need to know how much aluminium oxide is required. One mole of Al2O3 is needed to produce two moles of Al. So we need to divide the number of moles of Aluminium by 2 (1 mole of Al2O3 makes 2 moles of Al, so the number of moles of Al must be double that of Al2O3).

Amount of Al2O3 required = 20 = **10mol**

2

Finally, we can rearrange the mole equation to calculate the mass of Al2O3 needed to produce 540g of Al.

Mass of Al2O3 = Amount (mol) of Al2O3 x Mr

10mol x ((27 x 2) + (16 x 3)) = Mass of Al2O3

10 x 102 = **1020g**

**You will need to know these steps for your exam.**

**Balanced equations- continued**

This is for **higher tier ONLY**

The numbers needed to balance an equation can be calculated from the masses of the reactants and the products using moles.

In a chemical reaction, 72g of magnesium was reacted with exactly 48g of oxygen molecules to produce 120g of magnesium oxide.

We can use the number of moles of reactants and products to write a balanced equation for this reaction.

Amount (mol) = Mass

Atomic/formula mass

Amount of Mg = 72 = 3mol

24

Amount of O2 = 48 = 1.5mol

32

Amount of MgO = 120 = 3mol

40

We can put these moles into a symbol equation:

3Mg + 1.5O2 🡪 3MgO

We can simplify this to remove the decimal place in the equation.

2Mg + O2 🡪 2MgO

**Limiting reactants**

This is for **higher tier ONLY**

Sometimes when two chemicals react together, one chemical is completely used up during the reaction. When this happens, the reaction stops. This is called a ‘**limiting reactant’**. The other chemical, which is not completely used up, it is said to be in **excess**.

**Moles of a gas**

This is for **triple Science ONLY**

At room temperature and pressure, one mole of any gas always takes up a volume of 24dm3.

At normal room temperature and pressure (rtp):

Volume = amount (mol) x 24dm3

**Quick test!**

69g of sodium reacts with chlorine to produce sodium chloride:

2Na + Cl2 🡪 2NaCl

1. Calculate the number of moles of sodium present.

\_\_\_\_\_\_\_\_\_ mol

1. Calculate the number of moles of chlorine (Cl2) that would be required to react exactly with the sodium.

\_\_\_\_\_\_\_\_\_ mol

1. Calculate the mass of chlorine that would be required to react exactly with the sodium.

\_\_\_\_\_\_\_\_\_g

1. 300g of Calcium carbonate thermally decomposes into Calcium oxide and Carbon dioxide. Calculate the volume of carbon dioxide produced in the following reaction:

CaCO3 🡪 CaO + CO2

\_\_\_\_\_\_\_\_\_\_\_\_\_\_dm3

**Concentration**

This is for **higher tier ONLY**

The concentration of a solution is measured using units of mol/dm3.

Concentration of a solution = amount of substance (mol)

Volume (dm3)

If 1.00 mole of solute is dissolved to form a solution that has a volume of 1.00dm3, then the concentration of the solution will be 1.00mol/dm3.

2.00dm3 of sodium hydroxide solution contains 0.50mol of sodium hydroxide. Therefore, we need to use the above equation to calculate the concentration of the solution.

0.50mol = 0.25mol/dm3

2.00dm3

**Titrations**

This is for **higher tier ONLY**

Acids and alkalis react together to form a neutral solution. Titration is an accurate technique that can be used to find out how much of an acid is needed to neutralise an acid.

When neutralisation takes place, the hydrogen ions (H+) from the acid join with the hydroxide ions (OH-) from the alkali to form water (neutral pH).

H+(aq) + OH-(aq) 🡪 H2O(l)

You must use an indicator in titrations, such as universal indicator or phenolphthalein. Phenolphthalein turns alkali pink, but goes clear and colourless when neutralised, and is only used for strong acids and alkalis.

Hydrochloric acid, nitric acid and sulfuric acid are all examples of strong acids. Sodium hydroxide and potassium hydroxide are all examples of strong alkalis.

**Required practical – Determination of the reacting volumes of solutions of a strong acid and a strong alkali by titration.**

This is for **higher tier ONLY**

|  |  |
| --- | --- |
| **Sample method**   1. Wash and rinse a pipette with the alkali being used. 2. Use the pipette to measure out a known and accurate volume of the alkali. 3. Place the alkali in a clean, dry conical flask. 4. Add a suitable indicator, e.g. phenolphthalein. 5. Place the flask on a white tile so the colour can be seen clearly. 6. Place the acid into a burette that has been carefully washed with the acid. 7. Take a reading of the volume of acid in the burette (initial reading). 8. Carefully drip the acid to the alkali, swirling the flask to thoroughly mix. 9. Continue until the indicator just changes colour. This is called the end point. 10. Take a reading of the volume of acid in the burette (final reading). 11. Calculate the volume of acid added (i.e. subtract the final reading from the initial reading.) | **Hazards and risks**   * Acids and alkalis can damage the skin and eyes, so eye protection must be worn and any spillages wiped up |

Titration can be used to find the concentration of an acid or alkali, providing the following are known:

* The relative volumes of acid and alkali used.
* The concentration of the other acid or alkali.

Break down the calculation:

1. Write down a balanced equation for the reaction

to determine the ratio of moles of acid to alkali

involved.

1. Calculate the number of moles in the solution of

known volume and concentration. You can work out

the number of moles in the other solution from the balanced equation.

1. Calculate the concentration of the other solution.

For example, a titration is carried out and 0.04dm3 hydrochloric acid neutralises 0.08dm3 sodium hydroxide of concentration 1.00mol/dm3.

Balanced symbol equation: HCl + NaOH 🡪 NaCl + H2O

Number of moles of NaOH = volume x concentration

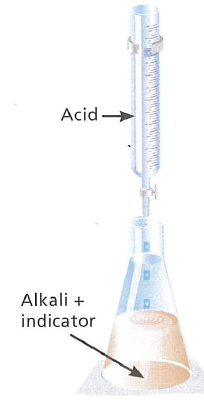
= 0.08dm3 x 1.00mol/dm3 = 0.08mol

Concentration of HCl = Number of moles of HCl

Volume of HCl

= 0.08mol = 2.00mol/dm3

0.04dm3



**Quick test!**

1. 1.50 moles of a solute dissolved in 1.00dm3 of solution. Calculate the concentration of the solution.

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ mol/dm3

1. 2.00g of solute was dissolved in 2.00dm3 of solution. Calculate the concentration of the solution in g/dm3.

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ g/dm3

1. A student wanted to calculate the concentration of sulfuric acid (H2SO4) by reacting it with 0.1dm3 of 2mol/dm3 sodium hydroxide (NaOH). A total of 0.05dm3 of H2SO4 was titrated. The products are sodium sulfate (Na2SO4) and water (H2O). Write the balanced symbol equation, and calculate the concentration of the acid.

Balanced symbol equation\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ml/dm3

**Percentage yield**

This is for **triple Science ONLY**

Atoms are never lost or gained in a chemical reaction. However, it is not always possible to obtain the calculated amount of product:

* If the reaction is reversible, it might not go to completion.
* Some product could be lost when it is separated from the reaction mixture.
* Some of the reactants may react in different ways to the expected amount.

The amount of product obtained is called the yield.

The percentage yield can be calculated using the formula:

Percentage yield = yield from reaction x 100

Maximum theoretical yield

**Calculating yield**

This is for **triple Science ONLY**

An example question could be: “*calculate how much calcium oxide can be produced from 50.0kg of calcium carbonate*.”

CaCO3 🡪 CaO + CO2

You need to work out the relative masses (Mr) of each compound.

(40 + 12 + (16x3)) 🡪 (40 + 16) + (12 + (16x2))

100 🡪 56 + 44

This tells you that 100kg of CaCO3 produces 56kg of CaO and 44kg of CO2.

So 1kg of CaCO3 produces 56 = 0.56kg of CaO.

100

Therefore, to turn 1kg of CaCO3 into 50kg of CaCO3, you multiply by 50. You can also multiply the mass of CaO by 50, as it is proportional.

50kg of CaCO3 produces 0.56 x 50 = 28kg of CaO.

A follow up question could be to calculate the actual percentage yield: “*A company heats 50kg of calcium carbonate in a kiln and obtains 22kg of calcium oxide. Calculate the percentage yield.”*

Percentage yield = 22 x 100 = 78.6%

28

**Atom economy**

This is for **triple Science ONLY**

Atom economy is a measure of the amount of reactant that ends up in a useful product. Scientists try to choose reaction pathways that have the highest atom economy. This is important for economic reasons and for suitable development, as more products are made and less waste is produced.

The percentage atom economy is calculated using the formula:

Atom economy = relative formula mass of the desired product x 100

Sum of the relative formula mass of all the

reactants

**The production of ethanol**

This is for **triple Science ONLY**

Ethanol can be produced in two different ways: hydration and fermentation.

During hydration, ethene is reacted with stea, to form ethanol:

C2H4 + H2O 🡪 C2H5OH

The atom economy for this method is 100%, as all of the reactants has turned into our desired product: we don’t need to calculate anything here! We call this an addition reaction.

However, with fermentation, not all of the reactants form the desired product. Ethanol is produced from the fermentation of glucose.

C6H12O6 🡪 2C2H5OH + 2CO2

We can use the atom economy equation to calculate how much glucose goes into ethanol.

Mr 2C2H5OH 🡪 2((12x2) + (1x5) + 16 + 1)) = 2 x 46 = 92

Mr C6H12O6 🡪 (12x6) + (1x12) + (16x6) = 72 + 12 + 96 = 180

Atom economy = 92 x 100 = 51.1%

180

The atom economy for this reaction is much lower - only about half of the atoms in the reactant turns into the desired product.

**Choosing a reaction pathway**

This is for **triple Science ONLY**

Comparing the atom economy of two competing reactions is important. However, it is just one of the factors that scientists have to consider when they choose which method to use. Important factors to consider when choosing a reaction pathway (and the units they can be found in, if relevant) include:

* The atom economy (Unit C3)
* The cost of reactants
* The percentage yield (Unit C3)
* The rate of reaction (Unit C6)
* The equilibrium position (Unit C6)
* The usefulness of by-products

**Quick quiz!**

1. Why might the actual yield of a reaction be less than the theoretical yield of the reaction?

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1. A reaction has a theoretical yield of 13g but an actual yield of 8.5g. calculate the percentage yield of this reaction.

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1. Explain why reactions that have a high atom economy good for the environment.

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1. The equation for the formation of ammonia is “N2 + 3H2 🡪 2NH3”. What is the percentage atom economy for this reaction?

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**Practice exam questions**

This is **FOR ALL ABILITIES**

1. Sulfur dioxide is produced when sulfur is burned. Ar  of sulfur and oxygen is S = 31, O = 16.

Calculate the relative formula mass (Mr) of SO2. (2 marks)

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. What is the Mr of glucose? C6H12O6. (2 marks)

Ar: C = 12, H = 1, O = 16

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. What is the Mr of Ca(NO3)2? (2 marks)

Ar: Ca = 40, N = 14, O = 16

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1. A student adds a piece of magnesium ribbon to a flask of dilute hydrochloric acid. The balanced symbol equation is as follow:

Mg(s) + 2HCl(aq) 🡪 MgCl2(aq) + H2(g)

1. Identify the definition of ‘the conservation of mass’ (1 mark).

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1. Explain why the mass of the reaction decreases as the reaction carries on (2 marks).

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1. Suggest a way to prevent the mass from being reduced during this reaction (1 mark).

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**Practice exam questions**

This is **FOR HIGHER TIER ONLY**

1. The avogardro constant has a value of 6.02x1023.
2. How many atoms are present in 7g of lithium?

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. How many atoms are present in 24g of carbon?

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. Calculate the number of moles in:
2. 19 of fluorine, F2

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. 22g of carbon dioxide, CO2

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. 17g of hydroxide, OH- ions.

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1. Complete combustion of carbon produces carbon dioxide, CO2

C + O2 🡪 CO2

1.8g of carbon was completely burned in oxygen.

1. How many moles of carbon were burned?

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. Calculate the mass of CO2 produced in this reaction.

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Practice exam questions**

This is **FOR HIGHER TIER ONLY**

1. Potassium iodide reacts with lead nitrate to produce potassium nitrate and lead iodide.

2KI(aq) + Pb(NO3)2(aq) 🡪 2KNO3(aq) + PbI2(s)

Write the ionic equation for this reaction (3 marks)

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1. When a hydrogen gas balloon explodes, the hydrogen reacts with an excess of oxygen to produce water vapour.

2H2 + O2 🡪 2H2O

1.8g of water was produced in this reaction.

1. What does the term ‘excess’ mean? (1 mark)

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1. Calculate the amount, in moles, of water vapour produced in this reaction. (2 marks)

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. Calculate the amount, in moles, of hydrogen that reacted in this reaction. (1 mark)

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. Calculate the mass of hydrogen that would produce 1.8g of water vapour. (2 marks)

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**Practice exam questions**

This is **FOR HIGHER TIER ONLY**

1. A student makes a solution of copper sulfate. They place 0.100 mole of copper sulfate crystals in a volumetric flask. They then add distilled water until the solution has a volume of 1.00dm3.

What is the concentration of this solution? You must give the correct unit. (2 marks)

\_\_\_\_\_\_\_\_\_\_\_\_ Unit: \_\_\_\_\_\_\_\_

1. Titration can be used to measure how much alkali is needed to neutralise an acid.

26.0cm3 of potassium hydroxide was placed in a flask.

The potassium hydroxide has a concentration of 0.2mol/dm3

This required 18.0cm3 of nitric acid solution for complete neutralisation.

The equation for the reaction can be summed up by the equation:

HNO3 + KOH 🡪 KNO3

1. How many moles of potassium hydroxide were used in this reaction? (2 marks)

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. How many moles of nitric acid were used in this reaction? (1 mark)

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. Calculate the concentration of the nitric acid. (2 marks)

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**Practice exam questions**

This is **FOR TRIPLE SCIENCE ONLY**

1. A student wanted to produce a sample of hydrogen gas by reacting either magnesium (Mg) or Zinc (Zn) with hydrochloric acid (HCl). The equations for these reactions are:

Mg + 2HCl 🡪 MgCl2 + H2

Zn + 2HCl 🡪ZnCl2 + H2

Determine which reaction is more economical by calculating and comparing the atom economy of both reactions. Use your periodic tables to find the relative atomic mass of each. (5 marks)

Atom economy of reaction using Mg = \_\_\_\_\_\_\_\_

Atom economy of reaction using Zinc = \_\_\_\_\_\_\_\_

1. Ammonia is produced in the Haber process by reacting nitrogen gas with hydrogen has. The equation for this reaction is:

N2 + 3H2 🡪 2NH3

1. A factory wanted to produce 14kg of nitrogen gas to ammonia. Calculate the theoretical yield of ammonia in this reaction. (2 marks)

Theoretical yield = \_\_\_\_\_\_\_\_\_\_\_\_ kg

1. After the reaction, the factory had produced 4.5kg of ammonia. Calculate the percentage yield for this reaction.(2 marks)

Percentage yield = \_\_\_\_\_\_\_\_\_\_\_\_ %

1. Give two reasons why it is desirable for a factory to obtain as high a percentage yield as possible (2 marks).

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