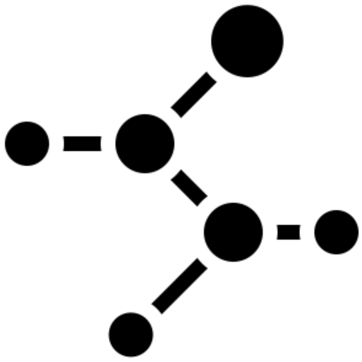
Quantitative



Chemistry

Name ________________________________

Class ________________________________

Teacher ________________________________

L1 Relative formula mass

What is Relative Formula Mass?

In simple terms, relative formula mass (RFM) is the sum of the atomic masses of all the atoms in a chemical formula. It tells us the mass of one molecule or formula unit of a compound. To calculate RFM, we need to know the atomic masses of the elements involved.

Calculating RFM Step-by-Step:

1. **Write Down the Formula:** Start by writing the chemical formula of the compound you want to find the RFM for. For example, let's take water (H_2O).
2. **Determine the Atomic Mass:** Find the atomic mass of each element in the formula. For hydrogen (H), it's approximately 1, and for oxygen (O), it's around 16.
3. **Multiply by the Subscript:** Multiply each atomic mass by the number of atoms of that element in the formula. In water, there are 2 hydrogen atoms and 1 oxygen atom.
4. **Add Up:** Add all the results together. In our example, it's 2(1) + 1(16) = 2 + 16 = 18.

So, the RFM of water (H_2O) is 18.

Why is RFM Important?

RFM is crucial in various aspects of chemistry. It helps us:

1. Determine the mass of reactants and products in chemical reactions.
2. Balance chemical equations.
3. Calculate the percentage composition of compounds.
4. Understand the stoichiometry of reactions.

Worked example 1:

What is the Mr of $CaCO_3$?

$CaCO_3$ has one calcium atom, one carbon atom and three oxygen atoms. The relative mass of calcium is 40, carbon is 12 and oxygen is 16. To calculate the total:

40 + 12 + (3 x 16) = 100

Worked example 2:

What is the Mr of $Ca(OH)_2$?

$Ca(OH)_2$ has one atom of calcium, two atoms of oxygen and two atoms of hydrogen (remember that everything in the brackets is multiplied by the little number. To calculate the total:

40 + (2 x 16) + (2 x 1) = 74

Independent practice

1. What is relative formula mass, and why is it important in chemistry?
2. **Extended writing (paragraph required):** Explain the steps to calculate the RFM of a compound.
3. What is the relative formula mass of water H_2O?
4. Calculate the relative formula mass of carbon dioxide CO_2.
5. What is the relative formula mass of methane CH_4?
6. Determine the relative formula mass of sodium chloride NaCl.
7. Calculate the relative formula mass of sulfuric acid H_2SO_4.
8. What is the relative formula mass of calcium carbonate $CaCO_3$?
9. Determine the relative formula mass of ammonia NH_3.
10. Calculate the relative formula mass of magnesium sulphate $MgSO_4$.
11. What is the relative formula mass of hydrogen peroxide H_2O_2?
12. Determine the relative formula mass of potassium nitrate KNO_3.

L2 Balancing equations.

In chemical reactions, the mass is conserved. This means that the mass of the reactants = the mass of the products. Sometimes it can look like the mass is changing, but it is usually down to a gas being involved:
If the mass appears to decrease, it is because a gas is formed which escapes the container.
If the mass appears to increase, it is because atoms from a gas in the atmosphere have been added.

<u>What is a Chemical Equation?</u>
Imagine chemical equations as recipes for reactions. They tell you what ingredients (reactants) go into a reaction and what you get as a result (products). But, there's a catch! For the recipe to work, you must have the same number of atoms on both sides of the equation.

<u>Why Balance Chemical Equations?</u>
Balancing equations is like making sure the recipe tastes just right. Unbalanced equations won't give you an accurate picture of what's happening in a reaction. To balance an equation, you adjust the coefficients (the numbers in front of chemical formulas) to make sure you have the same number of atoms on each side.

<u>The Steps to Balancing</u>
Let's take a simple example: the combustion of methane (CH_4) in oxygen (O_2) to form carbon dioxide (CO_2) and water (H_2O).

Step 1: Write down the equation: *$CH_4 + O_2$ -> $CO_2 + H_2O$*

Step 2: Count the atoms on both sides:

Carbon (C): 1 on the left, 1 on the right.

Hydrogen (H): 4 on the left, 2 on the right.

Oxygen (O): 2 on the left, 3 on the right.

Step 3: Adjust the coefficients. We start with carbon, so let's put a 1 in front of CO2:

$CH_4 + O_2$ -> $1CO_2 + H_2O$

Step 4: Re-count the atoms:

Carbon (C): 1 on the left, 1 on the right.

Hydrogen (H): 4 on the left, 2 on the right.

Oxygen (O): 2 on the left, 2 on the right (1 from CO_2 and 1 from H_2O).

Step 5: Balance hydrogen by putting a 2 in front of H_2O:

$CH_4 + O_2$ -> $1CO_2 + 2H_2O$

Step 6: Re-count the atoms:

Carbon (C): 1 on the left, 1 on the right.

Hydrogen (H): 4 on the left, 4 on the right.

Oxygen (O): 2 on the left, 4 on the right (2 from CO_2 and 2 from H_2O).

Step 7: Balance oxygen by putting a 2 in front of O_2:

$CH_4 + 2O_2$ -> $1CO_2 + 2H_2O$

Now, the equation is balanced with equal numbers of atoms on both sides.

Independent practice

1. What is the law of conservation of mass?
2. Why do we need to balance chemical equations?
3. What is the difference between reactants and products in a chemical equation?
4. What are coefficients in a chemical equation?
5. Describe all the steps in balancing a chemical reaction?
6. In the equation $C_6H_{12}O_6 + O_2 \text{->} CO_2 + H_2O$, how many carbon atoms are on each side after balancing?
7. Balance the equation: $N_2 + H_2 \text{->} NH_3$.
8. Balance the equation: $Fe + O_2 \text{->} Fe_2O_3$.
9. Balance the equation: $H_2 + O_2 \text{->} H_2O$
10. Balance the equation: $C + O_2 \text{->} CO_2$
11. Balance the equation: $Mg + HCl \text{->} MgCl_2 + H_2$
12. What does it mean if a chemical equation is unbalanced?
13. Why is it important to follow the law of conservation of mass when balancing equations?
14. Explain how you would balance the equation: $CH_3OH + O_2 \text{->} CO_2 + H_2O$.
15. In the equation $2KClO_3 \text{->} 2KCl + 3O_2$, how many oxygen molecules are produced when 6 moles of $KClO_3$ react?
16. What is the coefficient of O_2 in the balanced equation for the combustion of propane (C_3H_8)?
    $C_3H_8 + O_2 \text{->} CO_2 + H_2O$ – unbalanced
17. Balance the equation: $C_4H_{10} + O_2 \text{->} CO_2 + H_2O$.
18. How would you balance the equation: $NaOH + H_3PO_4 \text{->} Na_3PO_4 + H_2O$?

L3 Percentage by mass

1. What is Percentage by Mass?

Percentage by mass is a way of expressing the amount of a particular element in a compound relative to the total mass of the compound. It tells us how much of the compound is made up of a specific element.

2. Relative Atomic Mass (Ar)

Every element in the periodic table has a relative atomic mass, often abbreviated as "Ar." This value represents the average mass of all the isotopes of that element relative to the mass of a carbon-12 atom, which is assigned a mass of exactly 12 units. For example, the Ar of hydrogen is approximately 1, while the Ar of oxygen is about 16.

3. Relative Formula Mass (Mr)

When we have a compound, we calculate its relative formula mass, denoted as "Mr." This is the sum of the relative atomic masses of all the atoms in the compound, according to its chemical formula. For instance, in water (H2O), the Mr would be the Ar of hydrogen (1) multiplied by 2, plus the Ar of oxygen (16), resulting in a Mr of 18.

4. Calculating Percentage by Mass

To calculate the percentage by mass of an element in a compound, we follow a simple formula:

To calculate the percentage by mass of an element in a compound, we follow a simple formula:

$$\text{Percentage by mass of element} = \frac{\text{Ar of element} \times \text{number of atoms of element}}{\text{Mr of compound}} \times 100\%$$

Let's break this down with an example:

Example: Finding Percentage by Mass in Water (H2O)

Identify the Ar of each element:

- Ar of hydrogen (H) = 1
- Ar of oxygen (O) = 16

Calculate the Mr of the compound:

- Mr of water (H2O) = (Ar of H × 2) + Ar of O = (1 × 2) + 16 = 18

Plug the values into the formula:

- Percentage by mass of hydrogen = $\frac{1 \times 2}{18} \times 100\% = \frac{2}{18} \times 100\%$
- Percentage by mass of oxygen = $\frac{16}{18} \times 100\%$

Solve for the percentages:

- Percentage by mass of hydrogen ≈ 11.11%
- Percentage by mass of oxygen ≈ 88.89%

1. How is the relative formula mass (Mr) of a compound calculated, and why is it necessary for determining the percentage by mass of each element in the compound?
2. Using the compound carbon dioxide (CO_2) as an example, describe step-by-step how to calculate the percentage by mass of carbon and oxygen in the compound.
3. Discuss the importance of understanding the composition of compounds in various fields such as medicine, environmental science, and materials engineering.
4. Why is it crucial to use the correct relative atomic masses when calculating the percentage by mass of elements in compounds? Provide an example to illustrate your answer.
5. How does the knowledge of calculating percentage by mass in compounds relate to everyday life situations, such as cooking or understanding nutrition labels on food products? Provide specific examples to support your explanation.

Calculation practice

1. Calculate the percentage by mass of oxygen in water (H_2O), given that the relative atomic mass of hydrogen is 1 and oxygen is 16.
2. Determine the percentage by mass of carbon in carbon dioxide (CO_2), knowing that the relative atomic mass of carbon is 12 and oxygen is 16.
3. Find the percentage by mass of nitrogenz in ammonia (NH_3), with the relative atomic mass of nitrogen being 14 and hydrogen being 1.
4. Calculate the percentage by mass of sulphur in sulfuric acid (H_2SO_4), given the relative atomic masses: hydrogen (1), sulphur (32), and oxygen (16).
5. Determine the percentage by mass of calcium in calcium carbonate ($CaCO_3$), knowing the relative atomic masses: calcium (40), carbon (12), and oxygen (16).
6. Find the percentage by mass of chlorine in sodium chloride (NaCl), with the relative atomic masses: sodium (23) and chlorine (35.5).
7. Calculate the percentage by mass of iron in iron(III) oxide (Fe_2O_3), given the relative atomic masses: iron (56) and oxygen (16)

L4 Mass changes

Chemical reactions are fascinating processes that involve the rearrangement of atoms to form new substances. As you delve deeper into your studies of AQA GCSE Combined Science, one important concept you'll encounter is the observation of changes in mass during chemical reactions. Sometimes, these changes may seem puzzling at first, but fear not! By understanding the principles of conservation of mass and applying the particle model, you can make sense of these apparent discrepancies.

Let's start by addressing the observation that some reactions appear to involve a change in mass. One common scenario is when a reactant or product involved in the reaction is a gas. Gases, unlike solids or liquids, tend to escape into the surroundings, making it seem like there's a change in mass when it's just a matter of accounting for all the substances involved.

Consider the reaction between a metal and oxygen. When a metal reacts with oxygen, it forms a metal oxide. However, the mass of the oxide produced is often greater than the mass of the metal initially used. This seeming increase in mass can be perplexing if one doesn't consider the mass of oxygen involved. Oxygen is a gas, and its mass isn't typically considered in simple experiments. So, when it combines with the metal to form an oxide, its mass contributes to the overall mass of the products. Therefore, what initially seems like a discrepancy in mass is simply a result of not accounting for the mass of the gas involved.

Another example is the thermal decomposition of metal carbonates. When metal carbonates are heated, they undergo a chemical reaction to produce metal oxides, carbon dioxide gas, and in some cases, other products. However, carbon dioxide, being a gas, escapes into the atmosphere during the reaction. As a result, the mass of the solid products (such as metal oxides) remaining in the reaction vessel appears to be less than the initial mass of the metal carbonate. Again, this is because the mass of the gas (carbon dioxide) is not considered in such observations.

Now, let's delve into how we can explain these observed changes in mass using the particle model and balanced symbol equations. The particle model helps us understand the behaviour of substances at the microscopic level, focusing on the arrangement, movement, and interactions of particles.

When we balance a chemical equation, we ensure that the number of atoms of each element is the same on both sides of the equation. This balancing accounts for the conservation of mass, a fundamental principle in chemistry which states that mass is neither created nor destroyed during a chemical reaction, only rearranged.

For instance, let's consider the reaction between magnesium (Mg) and oxygen (O_2) to form magnesium oxide (MgO):

$2Mg + O_2 \rightarrow 2MgO$

In this equation, the number of magnesium atoms and oxygen molecules is the same on both sides, indicating a balanced reaction. However, if we were to consider the masses of all substances involved, including the oxygen gas, we would find that the total mass before the reaction is equal to the total mass after the reaction. It's just that the mass of the oxygen gas needs to be included in our calculations.

Similarly, in the thermal decomposition of metal carbonates like calcium carbonate ($CaCO_3$):

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

In this reaction, calcium carbonate decomposes into calcium oxide and carbon dioxide gas. As mentioned earlier, the carbon dioxide gas escapes into the atmosphere, leading to a perceived decrease in mass of the reaction vessel. However, if we consider the masses of all substances involved and balance the equation, we find that mass is conserved overall.

In conclusion, understanding changes in mass during chemical reactions requires considering all substances involved, including gases whose masses might not be readily apparent. By applying the principles of conservation of mass and utilizing the particle model to interpret reactions at the microscopic level, you can confidently explain any observed changes in mass in non-enclosed systems. Keep practicing balancing equations and applying these concepts, and you'll soon master this aspect of chemistry!

Independent practice

1. Explain why the mass of a metal oxide produced in a reaction with oxygen is often greater than the mass of the metal initially used.
2. How does the escape of carbon dioxide gas during the thermal decomposition of metal carbonates contribute to the observed changes in mass in non-enclosed systems?
3. Describe how the particle model helps to explain changes in mass during chemical reactions involving gases.
4. Why is it important to balance chemical equations when considering changes in mass during reactions?
5. Discuss how the principle of conservation of mass applies to chemical reactions and why it's a fundamental concept in chemistry.
6. **Long Answer Question:** Consider the reaction between magnesium and oxygen to form magnesium oxide. Explain, with reference to the particle model and balanced symbol equation, why the mass of the oxide produced is greater than the mass of the magnesium initially used.

L5 Uncertainty

Representation of Distribution of Results:

Imagine you're conducting an experiment where you're measuring the mass change during a chemical reaction. Let's say you're investigating the reaction between vinegar (acetic acid) and baking soda (sodium bicarbonate) to produce carbon dioxide gas. Each time you perform the experiment, you measure the mass change on a balance. Now, if you were to repeat this experiment multiple times, you'd likely get slightly different results each time.

The distribution of these results can be represented graphically using a histogram or a frequency polygon. This graph shows the spread of your measurements. You may notice that the measurements cluster around a central value, which is typically the mean (average) of your data points. However, not all measurements will be the same due to the inherent variability mentioned earlier.

Making Estimations of Uncertainty:

Uncertainty refers to the range within which the true value of a measurement is likely to lie. To estimate uncertainty in your measurements, one common approach is to consider the range of values obtained from multiple trials. For instance, if you performed the vinegar and baking soda experiment five times and recorded mass changes of 4.2 g, 4.1 g, 4.3 g, 4.0 g, and 4.2 g, the range of your measurements would be from 4.0 g to 4.3 g.

Using the Range as a Measure of Uncertainty:

The range of a set of measurements about the mean is a simple measure of uncertainty. It gives us an idea of how much the measurements vary from each other. In the example above, the range would be 4.3 g - 4.0 g = 0.3 g. This means that the mass change observed in the experiment could reasonably be anywhere within this range.

Understanding the significance of the range helps us appreciate the degree of uncertainty associated with our measurements. A larger range indicates greater variability among the measurements, leading to higher uncertainty, while a smaller range suggests more consistency and lower uncertainty.

Practical Application:

Let's say you need to determine the mass of a product formed in a chemical reaction. By conducting the reaction multiple times and calculating the range of mass changes observed, you can estimate the uncertainty associated with your measurements. This uncertainty should be considered when reporting your results.

For instance, if your average mass change is 4.2 g with a range of ±0.3 g, you would report your result as "the mass of the product formed is 4.2 ± 0.3 g". This indicates to others that there is some uncertainty in your measurement, and the true value could be anywhere within the given range.

Conclusion: In summary, understanding uncertainty in chemical mass changes is essential for accurate scientific analysis. By representing the distribution of results and using measures like the range of measurements about the mean, we can estimate and communicate the uncertainty associated with our experiments. This helps ensure the reliability and validity of our findings, fostering a deeper understanding of the natural world through science.

1. What is uncertainty in scientific measurements, and why is it important to consider?
2. How can the distribution of results be represented graphically in scientific experiments?
3. Explain the concept of range as a measure of uncertainty, using an example from a chemical reaction.
4. Why is it necessary to take into account the uncertainty associated with measurements when reporting experimental results?

Long Answer Question: Explain the process of estimating uncertainty in chemical mass changes using the range of measurements about the mean. Provide a step-by-step example to illustrate your explanation.

Calculation Questions:

1. If the mass changes observed in an experiment are 8.5 g, 8.7 g, 8.6 g, 8.4 g, and 8.8 g, what is the range of measurements?
2. In a different experiment, the average mass change is 12.3 g, with a range of ±0.2 g. What is the maximum and minimum possible mass change observed?
3. If the range of measurements for a reaction is 0.6 g and the mean mass change is 5.2 g, what is the upper and lower bound of uncertainty in the measurement?
4. A student measures the mass change in a chemical reaction four times and obtains 6.1 g, 6.2 g, 6.0 g, and 6.3 g. What is the average mass change?
5. In another experiment, the mass changes observed are 3.8 g, 3.9 g, 4.0 g, 3.7 g, and 4.1 g. What is the range of measurements?
6. If the range of measurements for a reaction is 0.4 g and the mean mass change is 9.7 g, what is the percentage uncertainty in the measurement?

L6 The mole

In chemistry, a mole is a unit used to measure the amount of a substance. The symbol for the unit mole is "mol." Think of it as like how you use "dozen" to represent 12 items. Similarly, a mole represents a specific number of particles.

2. Mass of one mole of a substance:

The mass of one mole of a substance in grams is numerically equal to its relative formula mass. This means that if you know the formula of a substance, you can find out its molar mass by adding up the atomic masses of all the atoms in its formula.

For example, let's consider water (H2O). The molar mass of water is the sum of the atomic masses of two hydrogen atoms and one oxygen atom: 2(H) + 1(O) = 2*1 + 16 = 18 grams per mole.

3. Avogadro's constant:

One mole of a substance contains the same number of particles, whether they are atoms, molecules, ions, or any other type of particle. This number is known as Avogadro's constant, which is approximately 6.02 x 10^23 per mole. This constant helps us understand the scale at which chemical reactions occur.

4. The significance of the Avogadro constant:

Understanding the Avogadro constant is essential because it allows us to relate macroscopic quantities, such as mass, to microscopic quantities, such as the number of atoms or molecules. For example, if you have one mole of carbon atoms (C), you'll have the same number of carbon atoms as you would in one mole of carbon dioxide molecules (CO2).

5. Examples:

- **Carbon (C):** In one mole of carbon (C), there are 6.02 x 10^23 carbon atoms.
- **Carbon Dioxide (CO2):** In one mole of carbon dioxide (CO2), there are 6.02 x 10^23 molecules of CO2. Since each molecule of CO2 contains one carbon atom, you still have 6.02 x 10^23 carbon atoms.

6. Application in chemistry:

The concept of moles is vital in various aspects of chemistry, including stoichiometry, where we calculate the quantities of reactants and products in a chemical reaction. By knowing the number of moles involved, we can determine the amounts of substances needed or produced.

In conclusion, understanding moles in chemistry is crucial for various reasons. It allows us to quantify the amount of substances involved in reactions, relate macroscopic observations to microscopic particles, and make accurate calculations in chemistry. By grasping the concept of moles and Avogadro's constant, you'll unlock the door to a deeper understanding of chemical reactions and their underlying principles.

Independent Practice

1. Define what a mole is in chemistry and explain its significance in chemical measurements. (2-3 sentences)
2. What is the mass of one mole of carbon dioxide (CO2) molecules? Show your calculation. (1-2 sentences)
3. State Avogadro's constant and explain its importance in chemistry. (1-2 sentences)
4. Compare the number of atoms in one mole of hydrogen atoms (H) with the number of molecules in one mole of water (H2O). (1-2 sentences)

Long Answer Question:

Explain how the concept of moles and Avogadro's constant are used in stoichiometry to calculate the quantities of reactants and products in a chemical reaction. Provide an example to illustrate your explanation. (4-5 sentences)

Calculation Questions:

1. Calculate the mass of one mole of sodium chloride (NaCl). (Atomic masses: Na = 23 g/mol, Cl = 35.5 g/mol)
2. Determine the number of moles present in 36 grams of water (H_2O). (Molar mass of water = 18 g/mol)
3. How many molecules are there in 2 moles of carbon dioxide (CO_2)? (Avogadro's constant = 6.02 x 10^23 per mole)
4. What is the mass of 3 moles of oxygen gas (O_2)? (Molar mass of oxygen = 16 g/mol)
5. Calculate the number of atoms in 0.5 moles of magnesium (Mg). (Avogadro's constant = 6.02×10^{23} per mole)
6. If you have 4 moles of ammonia (NH_3), how many nitrogen atoms are present? (NH3 contains 1 nitrogen atom per molecule)

L7 Using moles.

To understand moles in reactions we must understand how to use ratios. On a youth camp, there must be a ratio of at least one leader to six children. This can be written as a ratio of 1:6

If there are 12 children, there must be at least 2 leaders. This is because in a ratio you can multiply or divide either side to get to a target number. If the target is 12 children, we must multiply 6 by 2. But we must also multiply the other side by 2. This can be represented as

x 2   1:6   x 2
2: 12

However, this becomes a bit more complicated when we use different numbers. If there are 100 children, how many leaders do we need?

The easiest way to do this is in two steps. First divide the ratio by the original target side (in this case 6). Then multiply by 100 to get to your target of 100:

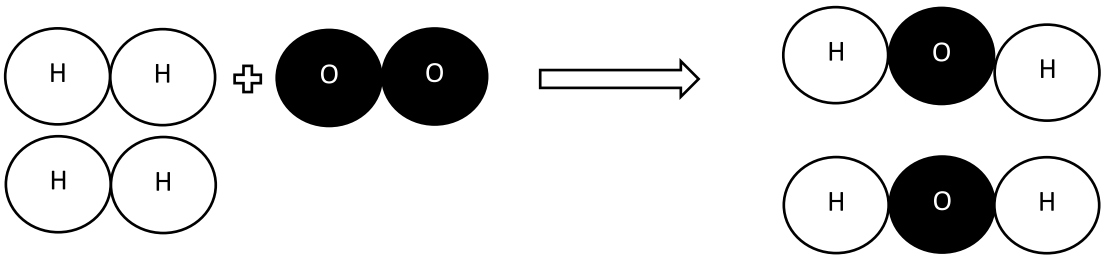
÷6   1:6   ÷6
0.17 : 1
x 100   17 : 100   x 100

This shows that if we have 100 children, we need 17 leaders.

This simple method can be applied to any ratio.

The above becomes relevant when we start to look at balanced equations. Take the equation as an example:

$2H_2 + O_2 \rightarrow 2H_2O$ can be shown as:



We can see that two molecules of hydrogen react with one molecule of oxygen to form two molecules of water. This can be expressed as a ratio:

H_2	O_2	H_2O
2	1	2

Using part 4 we could therefore establish that if we had four molecules of hydrogen and two of oxygen, we would obtain four of water. If I used a dozen hydrogen molecules, I would need half a dozen oxygen molecules and would obtain a dozen water molecules.

H_2	O_2	H_2O
A dozen	Half a dozen	A dozen

When we do a chemical reaction we do not use such tiny amounts, we use much larger amounts, which we can represent with the mole. If I have **two moles** of hydrogen I would need **one mole** of oxygen and would obtain **two moles of water** as below:

H_2	O_2	H_2O
Two moles	One mole	Two moles

If I was starting with 8 moles of hydrogen then it is obvious I would need 4 moles of oxygen. But it is more complicated if we use different numbers. If I started with 13.87 moles of hydrogen, how many moles of oxygen would I need?

	H_2	O_2	H_2O
	2	1	2
÷2	1	0.5	
x 13.87	13.87	6.94	

We can also now predict how much water we would expect to obtain from the reaction:

	H_2	O_2	H_2O
	2	1	2
÷2.	1	0.5	1
x 13.87	13.87	6.94	13.87

Worked example 4:

Hydrogen and nitrogen react together to make ammonia (NH_3). Write a balanced symbol equation for this reaction and calculate how much nitrogen would be needed to react with 19.30 moles of hydrogen and how much ammonia would be produced.

First, we write the equation:

$H_2 + N_2 \rightarrow NH_3$

Then balance:

$3H_2 + N_2 \rightarrow 2NH_3$

Then we calculate our ratio:

	H_2	N_2	NH_3
	3	1	2
÷3	1	0.33	0.67
x 19.30	19.30	6.37	12.93

So 6.37 moles of nitrogen would be needed and would produce 12.93 moles of ammonia.

Independent practice

1. Methane (CH_4) reacts with oxygen to make carbon dioxide and water.
    a. Write a balanced symbol equation for this reaction:
    CH_4 + __O_2 → CO_2 + __H_2O
    b. Fill in the top row of the table below with the numbers from your balanced equation

	CH_4	O_2	CO_2	H_2O
÷____				
x ____				

    c. If you started with two moles of CH_4, how much O_2 would you need? (the answer is 4, but you must use a calculation to prove this)
    d. If you started with 3.5 moles of CH_4, how much O_2 would you need?
    e. If you started with 3.5 moles of CH_4, how much CO_2 would you expect?
2. Ethane (C_2H_6) also reacts with oxygen to produce carbon dioxide and water, the equation being:
    $2C_2H_6 + 5O_2 \rightarrow 4CO_2 + 6H_2O$
    a. If 4 moles of ethane are used, show that 8 moles of CO_2 are produced
    b. If 5 moles of ethane are used, show that 15 moles of H_2O are produced
    c. If 7 moles of ethane are used, show that 21 moles of H_2o are produced.
    d. If 19 moles of oxygen is used, how much water is produced?
    e. How many moles of oxygen and ethane would you have to use to generate 43 moles of water?
3. Sulphur reacts with oxygen to make sulphur trioxide as below:
    $S_8 + O_2 \rightarrow SO_3$
    a. Balance the equation
    b. How many moles of sulphur would be required to produce 12 moles of sulphur trioxide?
    c. A chemist uses 17 moles of sulphur. How much oxygen would they need for a complete reaction?
    d. How much sulphur dioxide would be produced?

L8 Limiting reactants.

In chemistry, when we mix chemicals together to make something new, it's not always as straightforward as combining ingredients for a cake. Sometimes, we have to deal with something called **limiting reagents**, which can make things a bit trickier. But fear not! Once you understand the concept, it's not so scary.

Imagine you're baking cookies. You have 10 cups of flour, 6 eggs, and 4 cups of sugar. Each recipe calls for 2 cups of flour, 1 egg, and 1 cup of sugar. How many cookies can you make?
Here's where the concept of limiting reagents comes into play. The ingredient you have the least of, in this case, the eggs, limits how many cookies you can make. Even if you have plenty of flour and sugar left, you can only make 6 cookies because you don't have enough eggs. The eggs are the limiting reagent in this scenario.
In chemistry, it's similar. When we mix chemicals, one of them might run out before the others. That chemical is the limiting reagent because it determines how much product we can make.

How to Calculate Limiting Reagents
Let's break it down step by step with an example:
Example: Suppose we're mixing hydrogen gas (H2) with oxygen gas (O2) to make water (H2O). We have 4 moles of hydrogen and 3 moles of oxygen. How much water can we produce?
Step 1: Write down the balanced chemical equation.
The equation for the reaction between hydrogen and oxygen to form water is:
$2H_2 + O_2 \rightarrow 2H_2O$
Step 2: Calculate how much product each reactant can make individually.
Since the ratio of hydrogen to water is 2:2 and oxygen to water is 1:2, we need to calculate the amount of water each reactant can produce.
For hydrogen: 4 moles H2 * (2 moles H2O / 2 moles H2) = 4 moles H2O
For oxygen: 3 moles O2 * (2 moles H2O / 1 mole O2) = 6 moles H2O
Step 3: Identify the limiting reagent.
Compare the amounts of product each reactant can make. The reactant that produces the least amount of product is the limiting reagent. In this case, hydrogen produces 4 moles of water, while oxygen produces 6 moles of water. Therefore, hydrogen is the limiting reagent because it determines the maximum amount of water we can produce.
Step 4: Calculate the amount of product produced.
Since we're limited by hydrogen, we can produce 4 moles of water.
Practice Problem:
Now, let's try another example:

Example: We have 5 moles of nitrogen gas (N2) and 7 moles of hydrogen gas (H2). How much ammonia (NH3) can we produce?
Step 1: Write down the balanced chemical equation.
The equation for the reaction between nitrogen and hydrogen to form ammonia is:
$N_2 + 3H_2 \rightarrow 2NH_3$
Step 2: Calculate how much product each reactant can make individually.
For nitrogen: 5 moles N2 * (2 moles NH3 / 1 mole N2) = 10 moles NH3
For hydrogen: 7 moles H2 * (2 moles NH3 / 3 moles H2) ≈ 4.67 moles NH3
Step 3: Identify the limiting reagent.
Hydrogen produces the least amount of product, so it's the limiting reagent.
Step 4: Calculate the amount of product produced.
Since we're limited by hydrogen, we can produce approximately 4.67 moles of ammonia.

Understanding limiting reagents is crucial in chemistry because it helps us determine how much product we can make in a chemical reaction. By following these steps and practicing, you'll become a pro at calculating limiting reagents in no time!

Independent practice

1. What is a limiting reagent in a chemical reaction, and why is it important to identify?
2. In the example involving the reaction between hydrogen and oxygen to form water, how did we determine which reactant was the limiting reagent?
3. Why is it necessary to write down the balanced chemical equation before calculating limiting reagents?
4. Can you explain why it's important to know the amount of product each reactant can produce individually when determining the limiting reagent?

Extended Writing Question:

Imagine you are a chemist conducting an experiment to produce sulfuric acid. Describe the steps you would take to identify the limiting reagent in the reaction between sulphur trioxide and water to form sulphuric acid. Include explanations of how you would calculate the amounts of product each reactant can make and how you would determine the limiting reagent.

Practice Calculation Questions (Moles given):

1. If 3 moles of hydrogen gas (H_2) react with 4 moles of nitrogen gas (N_2) according to the equation $N_2 + 3H_2 \rightarrow 2NH_3$, which reactant is the limiting reagent?
2. Given 5 moles of carbon monoxide (CO) and 8 moles of hydrogen gas (H_2), which reactant is the limiting reagent in the reaction $CO + 3H_2 \rightarrow CH_4 + H_2O$?
3. If 10 moles of oxygen gas (O_2) react with 6 moles of hydrogen gas (H_2) according to the equation $2H_2 + O_2 \rightarrow 2H_2O$, which reactant is the limiting reagent?

Practice Calculation Questions (Masses given):

1. If 12 grams of sodium (Na) react with 16 grams of chlorine gas (Cl_2) to form sodium chloride (NaCl) according to the equation $2Na + Cl_2 \rightarrow 2NaCl$, which reactant is the limiting reagent?
2. Given 25 grams of phosphorus (P) and 15 grams of oxygen gas (O_2), which reactant is the limiting reagent in the reaction $4P + 5O_2 \rightarrow 2P_2O_5$?
3. If 30 grams of calcium carbonate ($CaCO_3$) react with 20 grams of hydrochloric acid (HCl) according to the equation $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$, which reactant is the limiting reagent?

L9 Concentrations

Chemical reactions often occur in solutions, where substances dissolve in a liquid to form a homogeneous mixture. The concentration of a solution is a crucial aspect to understand as it helps us quantify the amount of solute (the substance being dissolved) present in a given volume of solvent (the liquid in which the solute is dissolved). In this guide, we'll delve into how to calculate concentration and explore the relationship between the mass of solute, the volume of solution, and concentration.

What is Concentration?

Concentration measures the amount of solute dissolved in a given amount of solvent. It's typically expressed in units like grams per decimetre cubed (g/dm^3). This tells us how much of the solute is present in a specific volume of the solution.

Calculating Concentration:

To calculate the concentration of a solution, we use the formula:

Concentration=Mass of Solute ÷ Volume of Solution

Let's work through an example:

Example:

Suppose we have 50 grams of salt (NaCl) dissolved in 250 cm^3 of water. Calculate the concentration of the salt solution in grams per decimetre cubed (g/dm^3).

Solution:

1. Convert the volume from cm^3 to dm^3.

   $250\ cm^3 = 250 \div 1000\ dm^3 = 0.25\ dm^3$

2. Now, use the formula to calculate concentration.

   Concentration= Mass of Solute ÷ Volume of Solution

Concentration=$50\ g \div 0.25\ dm^3$

Concentration=$200\ g/dm^3$

So, the concentration of the salt solution is 200 grams per decimetre cubed (g/dm^3).

Relationship between Mass, Volume, and Concentration:

The concentration of a solution depends on both the mass of solute and the volume of the solution. Here's how they're related:

1. **Direct Relationship:** As the mass of solute increases while the volume of the solution remains constant, the concentration increases. This is because there's more solute present in the same volume of solvent.

2. **Inverse Relationship:** Conversely, if the volume of the solution increases while keeping the mass of solute constant, the concentration decreases. This is because the same amount of solute is now distributed in a larger volume of solvent, making it less concentrated.

Consider a scenario where you're making lemonade. If you add more sugar (solute) to the same amount of water (solvent), the lemonade becomes sweeter, indicating an increase in concentration. However, if you add the same amount of sugar to a larger volume of water, the lemonade tastes less sweet because the sugar is more diluted in the larger volume, resulting in a lower concentration.

Independent practice

1. What does the concentration of a solution measure, and how is it typically expressed?
2. Describe the relationship between the mass of solute and the concentration of a solution.
3. Explain the effect on concentration when the volume of a solution is increased while keeping the mass of solute constant.
4. Provide an example from daily life where understanding concentration in solutions is important.

Extended Writing Question (Describe):

Describe the relationship between the mass of solute, the volume of solution, and the concentration of a solution, providing examples to illustrate your explanation.

Practice Calculation Questions:

1. A solution contains 30 grams of salt dissolved in 500 cm^3 of water. Calculate the concentration of the salt solution in g/dm^3.
2. If 60 grams of sugar are dissolved in 250 cm^3 of water, what is the concentration of the sugar solution in g/dm^3?
3. Calculate the concentration of a solution containing 25 grams of potassium chloride dissolved in 0.2 litres of water.
4. A solution has a concentration of 180 g/dm^3. If 90 grams of solute are dissolved, what volume of solution is produced?
5. What is the concentration of a solution if 40 grams of solute are dissolved in 0.5 litres of solvent?
6. If a solution has a concentration of 250 g/dm^3 and a volume of 0.4 dm^3, how much solute is present in the solution?