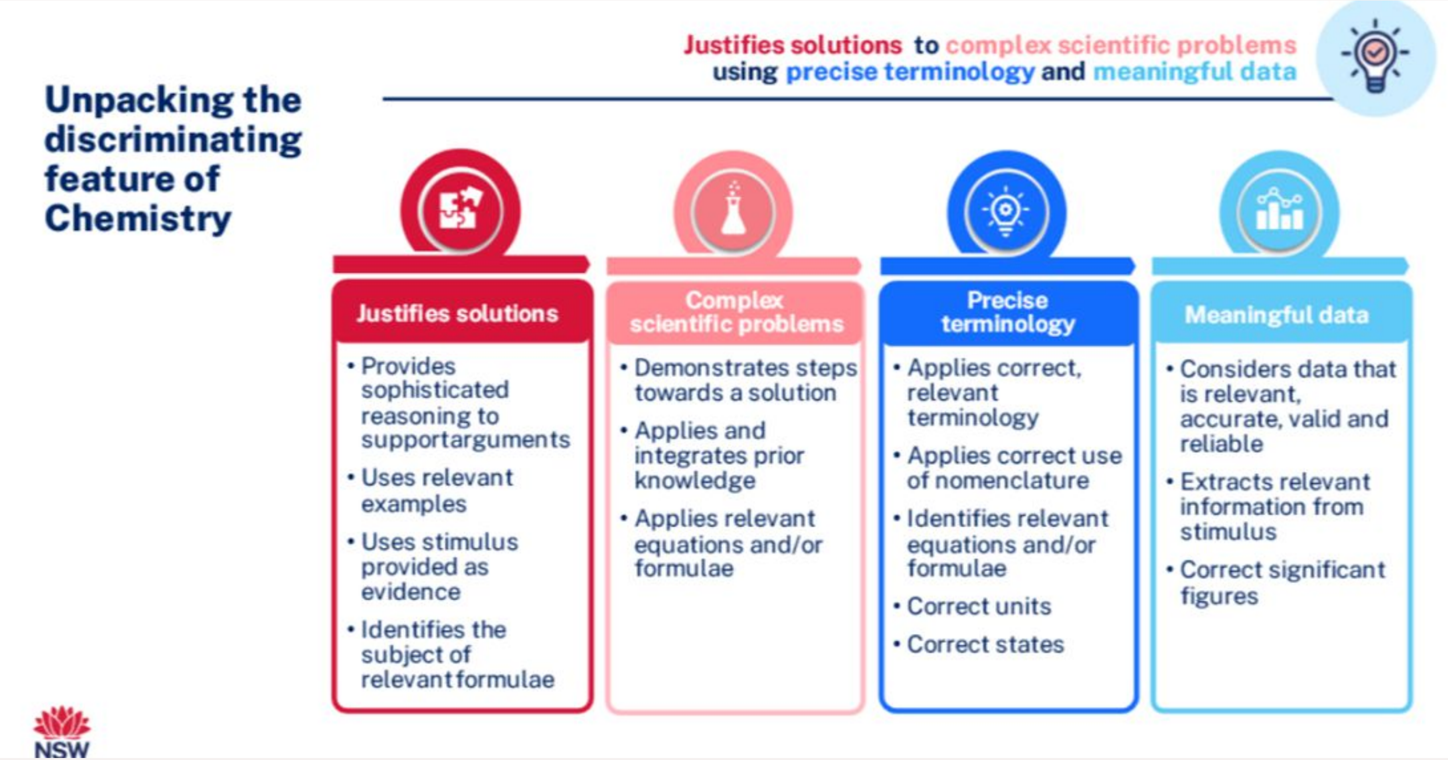
Year 11 Chemistry

Discriminating Features in Chemistry

Justifies solutions to complex scientific problems using precise terminology and meaningful data

Discriminating Features in Chemistry



Module 2

Introduction to Quantitative Chemistry

Inquiry Question 2:

How are measurements made in Chemistry?

2.1.1.a.b

- conduct practical investigations to observe and measure the quantitative relationships of chemical reactions, including but not limited to:

  – masses of solids and/or liquids in chemical reactions

  – volumes of gases in chemical reactions (ACSCH046)

Mass in Chemical Reactions

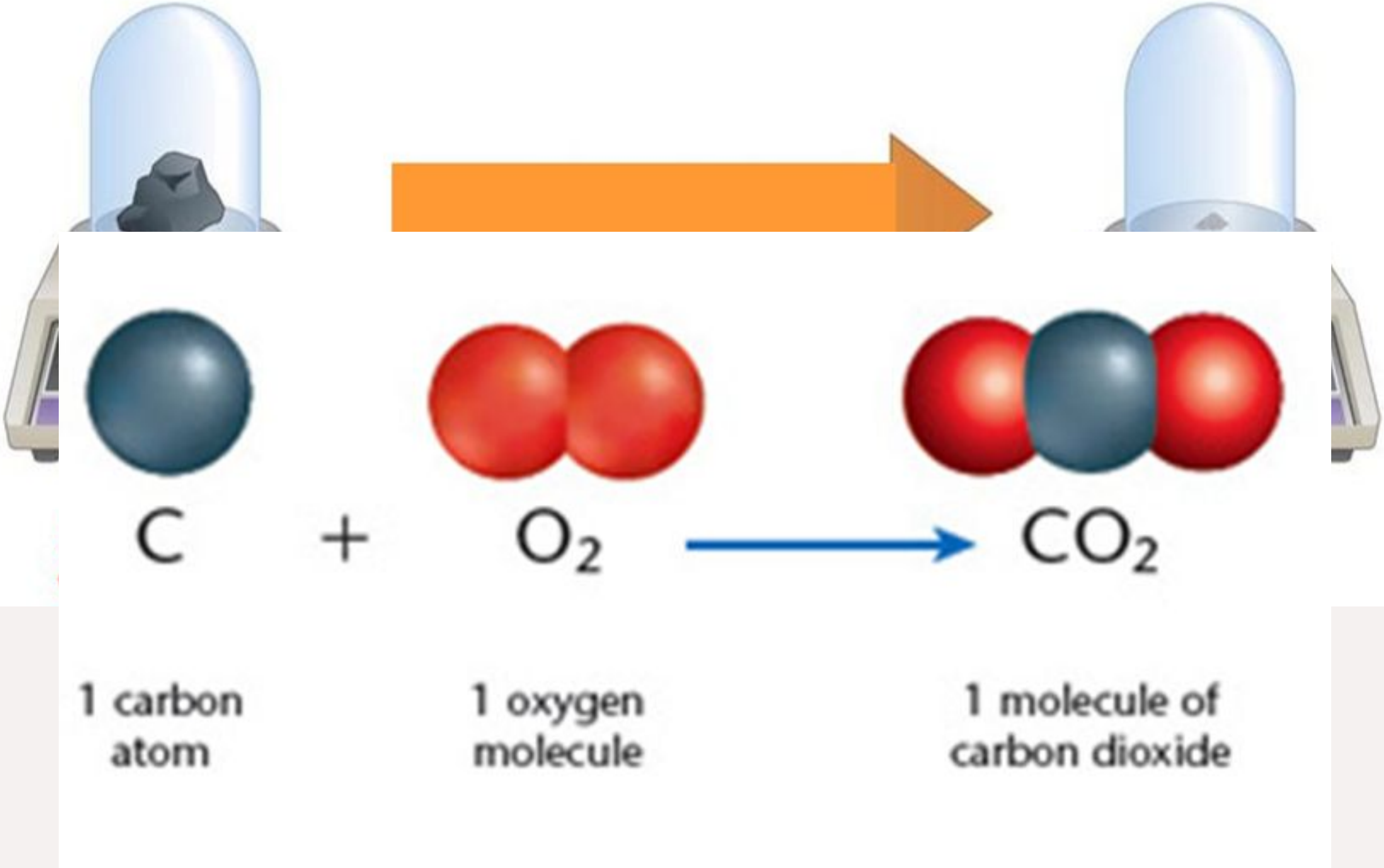
Practical Investigation

METHOD

1. Add 1 teaspoon of calcium chloride to the bag.
2. Add 1 teaspoon of sodium hydrogen carbonate to the bag.
3. Place 10 mL of water and four drops of universal indicator solution into the glass vial.
4. Carefully stand the glass vial in the bag so it doesn't fall over.
5. Seal the bag.
6. Weigh the sealed bag.
7. Invert the bag so the universal indicator solution mixes with the solids.
8. Record observations.
9. Reweigh the sealed bag.

ANALYSIS OF RESULTS

1. How could you tell that a chemical reaction was occurring?
2. What was the purpose of doing the investigation in a sealed bag?
3. Justify whether the results would have been the same if the investigation was not carried out in the sealed bag.


C
+
O_2
CO_2
1 carbon atom
1 oxygen molecule
1 molecule of carbon dioxide

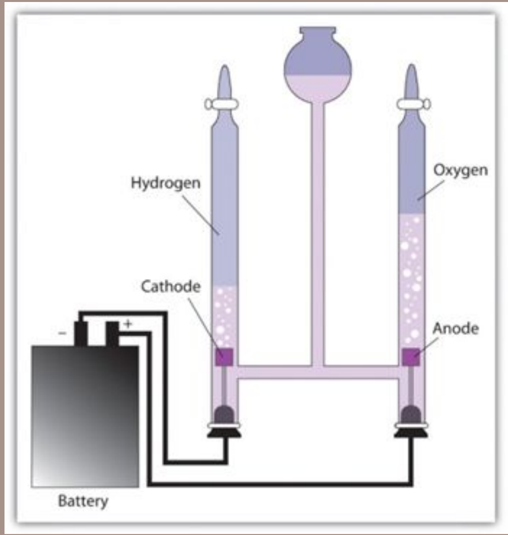
Practical Investigation

Volume in Chemical Reactions

Practical Investigation

Electrolysis of Water

	Initial Volume	Final Volume
Oxygen		
Hydrogen		



Inquiry Question 2:

How are measurements made in Chemistry?

2.1.2.a

- relate stoichiometry to the law of conservation of mass in chemical reactions by investigating:

  – balancing chemical equations (ACSCH039)

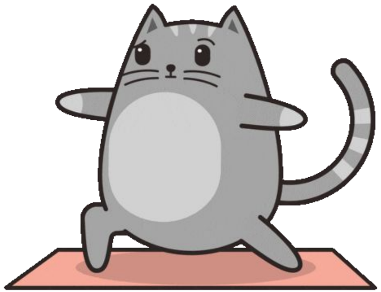
Stoichiometry

What is stoichiometry?

Stoichiometry is the study of quantitative relationships /ratios in chemical reactions.

- Involves the amounts of reactants used and of products produced

Balancing Equations



When balancing equations, the number of each atom has to be **equal on both sides**.

- It is easier to write down the number of atoms per element on each side (Remember to look out for brackets and subscripts)

- Add coefficients to balance the equation - leave Hydrogen and Oxygen till the end

Practise

makes perfect

1. $KBr_{(s)} + Cl_{2(g)} \rightarrow KCl_{(g)} + Br_{2(g)}$

2. $ZnS_{(s)} + O_{2(g)} \rightarrow ZnO_{(g)} + SO_{2(g)}$

3. $C_3H_{8(s)} + O_{2(g)} \rightarrow H_2O_{(l)} + CO_{2(g)}$

4. $Pb(NO_3)_{2(aq)} + NaCl_{(s)} \rightarrow NaNO_{3(aq)} + PbCl_{2(aq)}$

TO DO LIST:

- ❑ Writing balanced equations
- ❑ Balancing Equations

Inquiry Question 2:

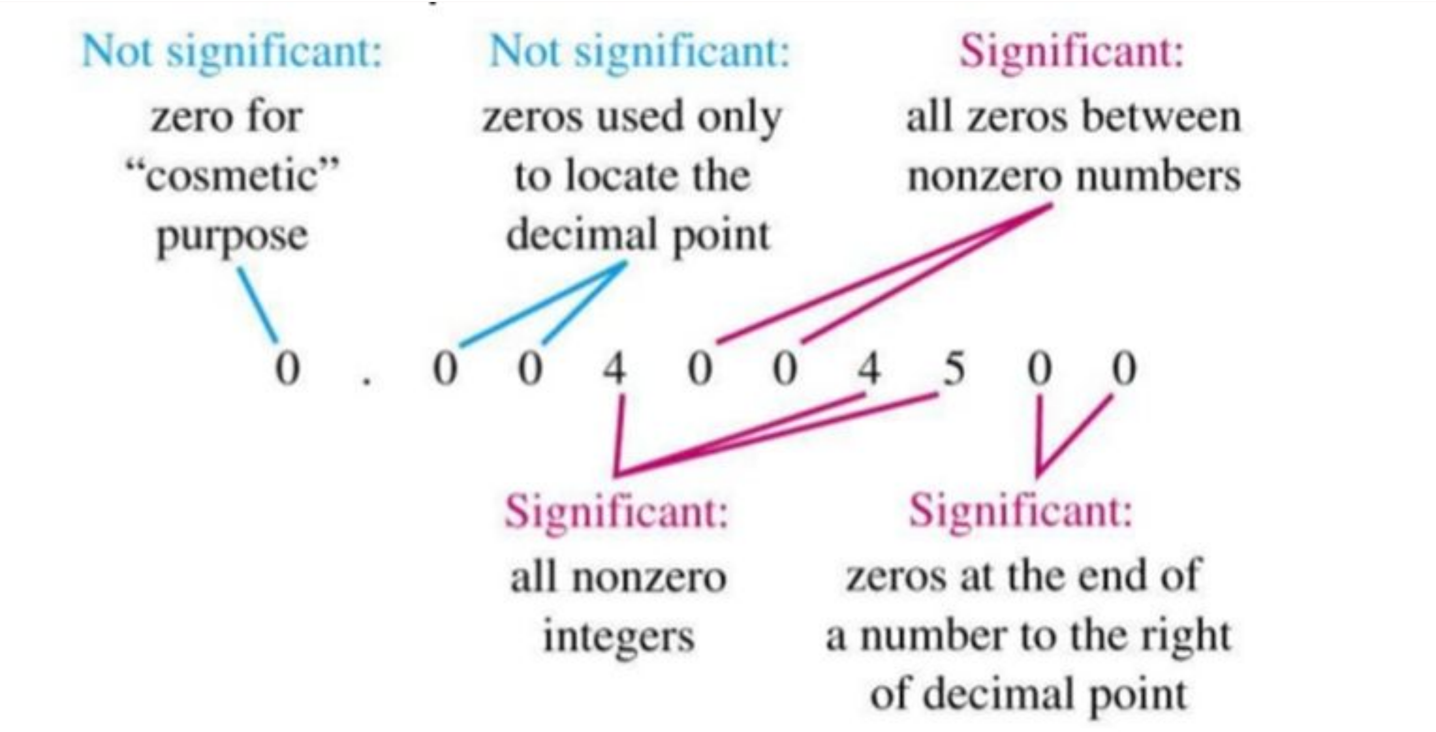
How are measurements made in Chemistry?

2.1.2.b

- relate stoichiometry to the law of conservation of mass in chemical reactions by investigating:

  – solving problems regarding mass changes in chemical reactions (ACSCH046)

Significant Figures & Scientific Notation



Trailing zeros in a number are significant only is the number contains a decimal point
E.g. 100.0 has 4 sig figs
100 has 1 sig fig

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makes perfect

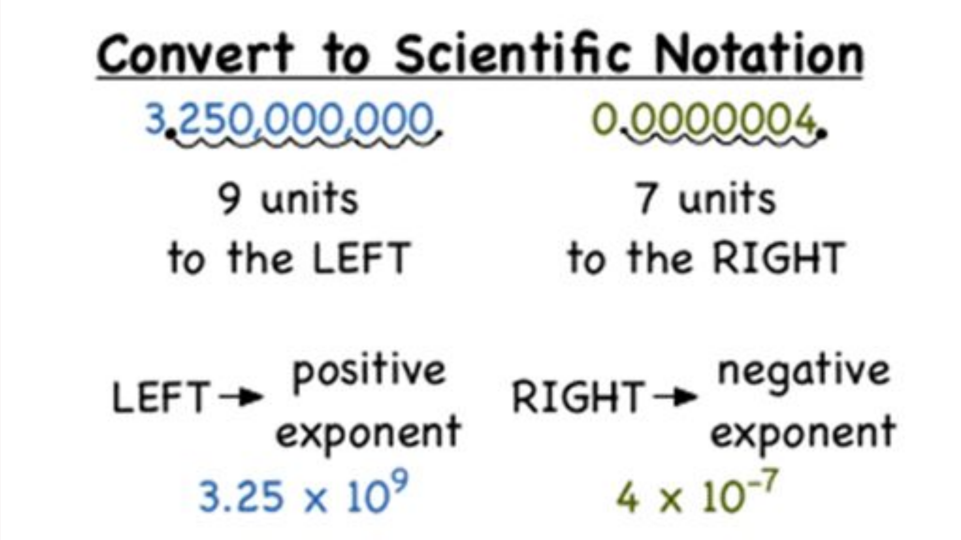
Determine the significant figures of each:

1. 0.0035
2. 1.080
3. 2371
4. 2.97×10^5
5. 10,001
6. 6,051.00
7. 0.1020
8. 0.00050

Significant Figures & Scientific Notation

Scientific Notation - used to express number that are very large or very small

- Always include all numbers that are not leading or trailing zeros



Practise

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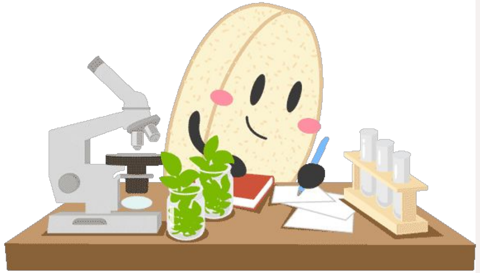
Convert the following to scientific notation

1. 0.005
2. 5050
3. 0.000008
4. 1,000,000
5. 0.00205
6. 0.0000870
7. 205412
8. 821654

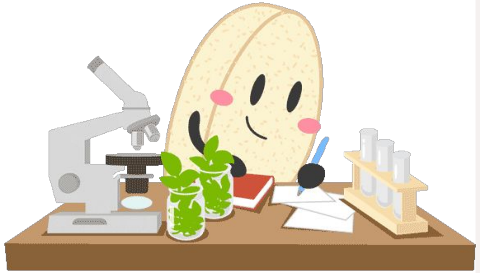
Mass - Mass Stoichiometry



Mass of product formed in a reaction or how much mass of reactant has been used up in a chemical reaction can be calculated.

Example 1:
Calculate the mass of product formed when 46.0g of sodium metal reacts with 253.8g of solid iodine, what mass of sodium iodide is produced?

Mass - Mass Stoichiometry with excess reactants



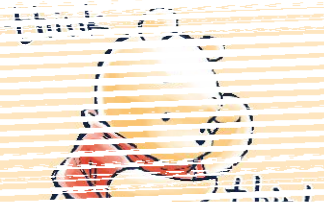
When a reactant is not all consumed in a reactant, it is known as the **excess**.

Example 2:

One students measures 34.5g of sodium reacts with 380.7g of iodine which is all consume and produces 415.2g of sodium iodide. Another student measures the same amount of sodium and reacts it with 450g of iodine and produces the same amount. What is the mass of iodine in excess?

Mass - Volume Stoichiometry

Why might this be the case?

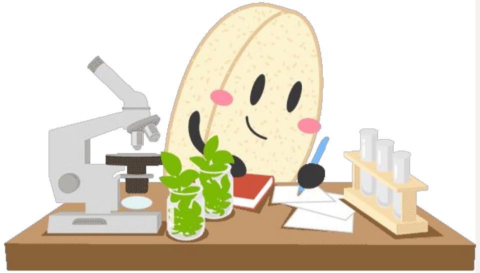


Volume and mass can be determined can be determined by knowing the density of the liquid.

$$d = m / v$$

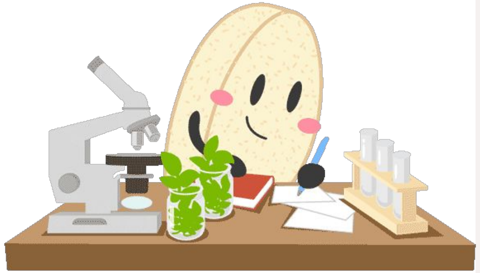
e.g. the density of water is 1g/mL at 4℃. However at 20℃ density of water is 0.998g/mL and 65℃ is 0.980g/mL.

Mass - Volume Stoichiometry



Example 3:
1.8kg of glucose from sugar cane is fermented in anaerobic conditions at 25℃ to produce 1.167L of ethanol (C_2H_5OH) and carbon dioxide. Determine the mass of ethanol and carbon dioxide produced. (Density of ethanol at 25℃ is 0.789g/mL).

Volume - Volume Stoichiometry



Example 4:

At 25℃ and 1 atm pressure, the density of methane (CH_4) is 0.656g/L, the density of oxygen is 1.31g/L and density of carbon dioxide is 1.81g/L/. If 12.2L of methane reacts completely with 24.4L of oxygen to produce 18g of water, what mass of carbon dioxide is produced?

Practise

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1. 0.04533L of hydrogen peroxide (H_2O_2) is produced from the reaction between 4.0g of hydrogen gas and sufficient oxygen to allow the reaction to be complete. Determine the mass of hydrogen peroxide produced and hence the mass of oxygen gas consumed.

2. 84.6g of solid iodine reacts with sufficient Na metal to produce 99.9g of sodium iodide, what mass of sodium is consumed?

3. A student use 11.5g of sodium and reacts it with 126.9g of iodine and produces 1338.4g of sodium iodide. Another student ues 15g of sodium and reacts it the same amount of iodine and gets the same amount produces. What mass of sodium is in excess?

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4. At 15℃ and 1 atm pressure, the density of butane (C_4H_{10}) is 2.48g/L, the density of oxygen gas is 1.36g/L and the density of carbon dioxide is 1.87g/L. Butane reacts with oxygen to form carbon dioxide and water. If 24.4L of butane reacts completely with 152.1L of oxygen to produce 90.0g of water, what mass of carbon dioxide is produced?

TO DO LIST:

- ❑ Textbook questions Page 217
- ❑ Textbook questions Page 219
- ❑ https://www.sciencegeek.net/Activities/GramsGramsStoich.html

Inquiry Question 2:

How are measurements made in Chemistry?

2.2.1

- conduct a practical investigation to demonstrate and calculate the molar mass (mass of one mole) of:

  – an element

  – a compound (ACSCH046)

Practical Investigation

Molar Mass of an Element and a Compound

Inquiry Question 2:

How are measurements made in Chemistry?

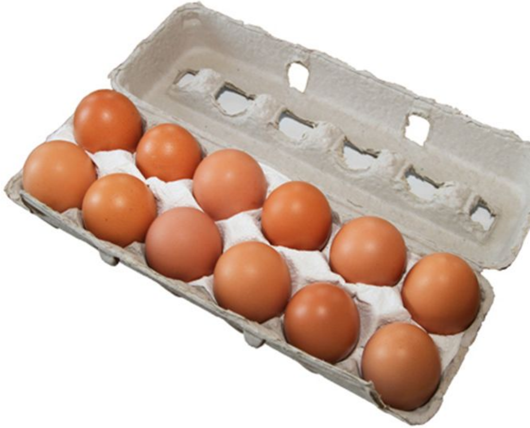
2.2.3.a

- explore the concept of the mole and relate this to Avogadro's constant to describe, calculate and manipulate masses, chemical amounts and numbers of particles in: (ACSCH007, ACSCH039)

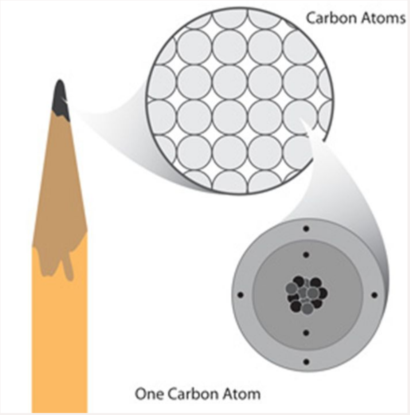
  – moles of elements and compounds $n=m/MM$ (n = chemical amount in moles, m = mass in grams, MM = molar mass in gmol-1)

Mole Concept

How would you describe the quantity of the following?



In Chemistry, atoms and molecules are extremely tiny!



The Mole

Chemists use mole to q

One mole consists of:

602, 214, 129, 000, 000

The Relationship between Moles, Particles and Mass

1 mole is equal to the atomic mass in grams.

E.g.

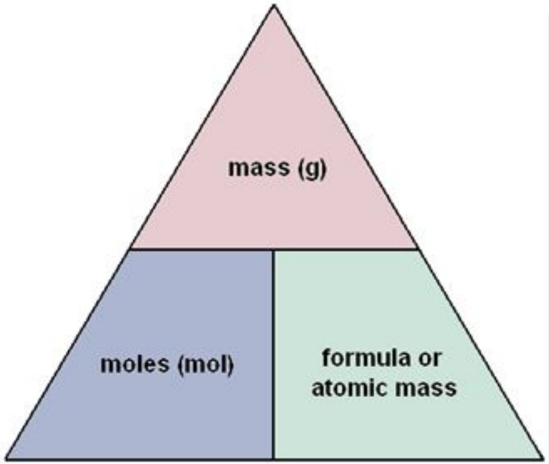
Molar mass - **mass of one mole** of a substance (g/mol)

- Molar mass of a compound is calculated by converting formula weight to g/mol

E.g. Water, Carbon dioxide, Glucose

The Relationship between Moles, Particles and Mass

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



Practise

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1. How many atoms in 14 moles of carbon?
2. How many moles in 4.00×10^{23} atoms of sulfur?
3. How many moles in 4.3×10^{22} molecules of H_3PO_4?
4. How many moles in 127.5g of NaCl?
5. What is the mass of 0.850 moles of sulfur dioxide?
6. Determine the number of moles in 32.7g of Ethanol (C2H6O).
7. Determine the mass of 4.3 moles of aluminium oxide.

RECAP

Answer the following questions:

1) How many moles are in 25.0 grams of water?

2) How many grams are in 4.500 moles of Li_2O?

3) How many molecules are in 23.0 moles of oxygen?

4) How many moles are in 3.4 x 10^{23} molecules of H_2SO_4?

5) How many molecules are in 25.0 grams of NH_3?

6) How many grams are in 8.200 x 10^{22} molecules of N_2I_6?

Inquiry Question 1:

What happens in chemical reactions?

2.2.2

- conduct an investigation to determine that chemicals react in simple whole number ratios by moles

Moles and Chemical Reactions

note:

The coefficient in a balanced equation will correspond to the number of moles.

E.g.

$$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

- This is known as **stoichiometry**

Practical Investigation

Combustion of Magnesium

Practical Write Up

You will be required to write up a proper scientific report on the Magnesium Practical.

Components:

- × Title
- × Aim
- × Hypothesis
- × Equipment
- × Risk Assessment (not Pre-Lab)
- × Method (in your own words - with correct format)
- × Results table
- × Discussion (reliability, validity, accuracy, errors/improvements)
- × Conclusion

Inquiry Question 2:

How are measurements made in Chemistry?

2.2.3.b

- explore the concept of the mole and relate this to Avogadro’s constant to describe, calculate and manipulate masses, chemical amounts and numbers of particles in: (ACSCH007, ACSCH039)

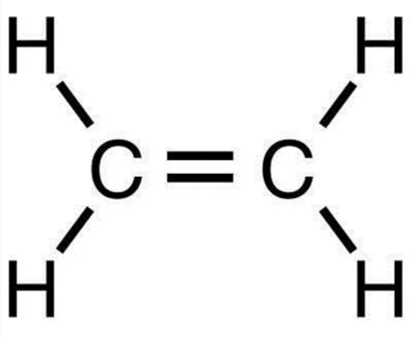
  – percentage composition calculations and empirical formulae

Molecular & Empirical Formula

Molecular formula shows the amount of atoms of each element in a compound.

Empirical formula is the **simplest reduced ratio** of atoms in a compound .

E.g.



Molecular formula

"regular formula"

How many atoms of each element are in a compound.

$$C_5H_{10}O_5$$

Ask yourself, what is the largest number that we divide each of these subscripts by to get to the empirical formula

Empirical formula

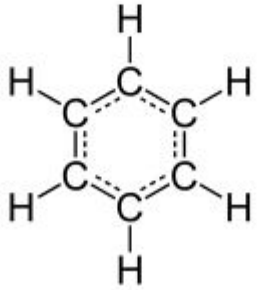
The simplest or most reduced ratio of atoms in a compound.

Molecular formula

"regular formula"

How many atoms of each element are in a compound.

Word formula:
Benzene



Empirical formula

The simplest or most reduced ratio of atoms in a compound.

Molecular formula

"regular formula"

How many atoms of each element are in a compound.

P_3N_5

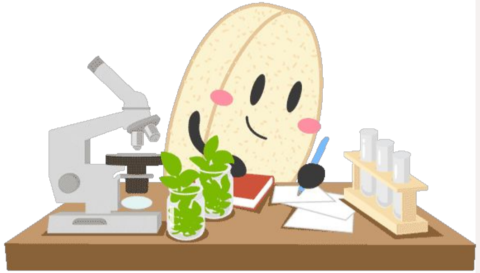
C_5H_{12}

Empirical formula

The simplest or most reduced ratio of atoms in a compound.

If the ratio of atoms in the molecular formula can't be simplified anymore, the empirical formula is the same as the molecular formula.

Calculating Molecular Formula from Empirical formula

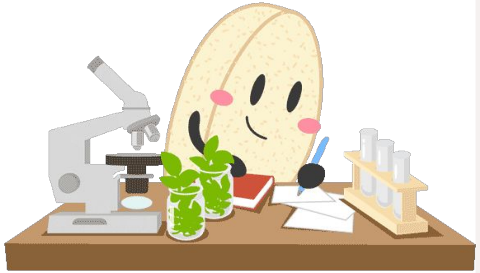


What we multiply by:

- $\frac{Molar\ mass\ of\ molecular\ formula}{Molar\ mass\ of\ empirical\ formula}$

1. Determine the molecular formula of a compound with the empirical formula CF_2 and a molar mass of 200.04 g/mol.

Calculating Molecular Formula from Empirical formula

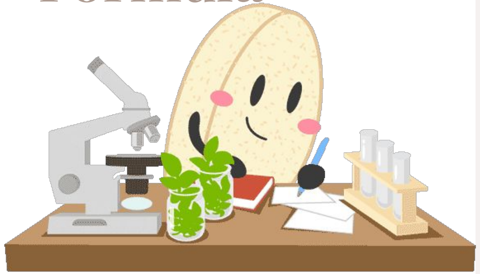


What we multiply by:

- $\frac{Molar\ mass\ of\ molecular\ formula}{Molar\ mass\ of\ empirical\ formula}$

2. A compound has empirical formula C_2H_5N and molar mass 86.16 g/mol. What is its molecular formula?

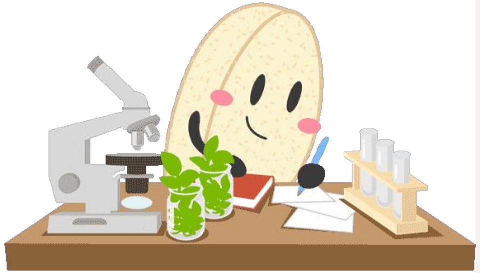
Calculating Percentage Composition from Empirical Formula



$$\frac{\text{molar mass of element}}{\text{molar mass of compound}} \times 100\%$$

3. Calculate the percentage composition of potassium and oxygen, which has an empirical formula of K_2O.

Calculating Empirical Formula from Percentage Composition



1. Divide each % by its molar mass

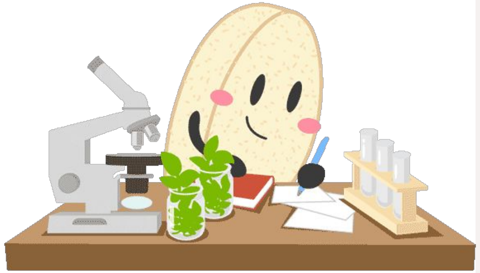
$$\frac{percentage\ compositon}{molar\ mass}$$

2. Divide each of those by the smallest number

3. Find the lowest whole number ratio

4. Find the empirical formula of a compound which contains 83% potassium and 17% oxygen.

Calculating Empirical Formula from Percentage Composition



1. Divide each % by its molar mass

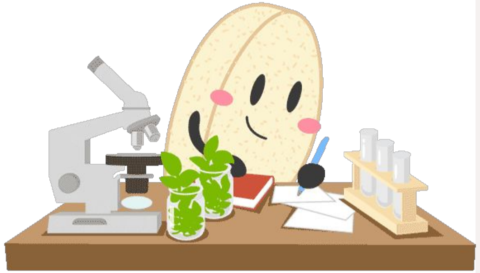
$$\frac{percentage\ compositon}{molar\ mass}$$

2. Divide each of those by the smallest number

3. Find the lowest whole number ratio

5. Calculate the empirical formula of a compound composed of 38.7% C, 16.2% H and 45.1% N.

Calculating Molecular Formula from Percentage Composition



1. Find its empirical formula
2. Calculate the molecular formula using:

$$\frac{molar\ mass\ of\ compound}{empirical\ molar\ mass}$$

6. The percentage composition of a hydrocarbon is 80% carbon and 20% hydrogen. Calculate its molecular formula if the molar mass of this compound is 30g/mol.

Inquiry Question 2:

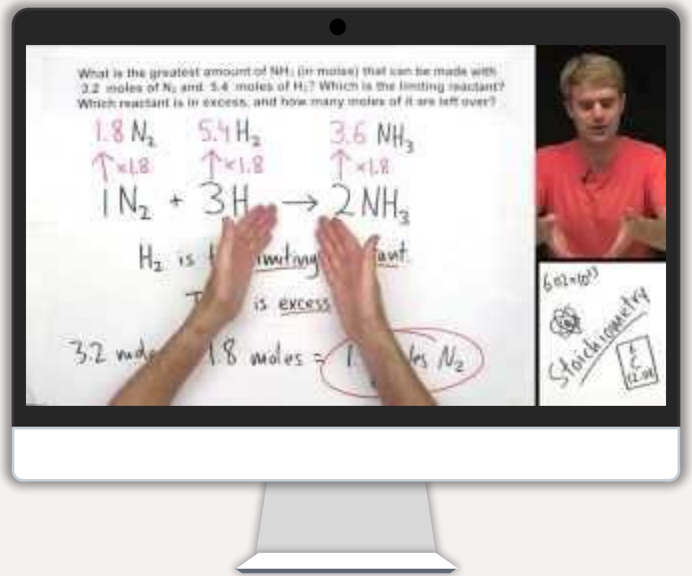
How are measurements made in Chemistry?

2.2.3.c

- explore the concept of the mole and relate this to Avogadro's constant to describe, calculate and manipulate masses, chemical amounts and numbers of particles in: (ACSCH007, ACSCH039)

  – limiting reagent reactions

Limiting Reagent and Excess Reactant



Practise

makes perfect

1. What is the greatest amount of NH_3 that can be made with 3.2 moles of N_2 and 5.4 moles of H_2?

$$N_2 + 3H_2 \rightarrow 2NH_3$$

What would be the limiting reactant?

Which reactant is in excess, and how many moles of it are left over?

2. What is the greatest amount of MgO (in moles) that can be made with 7.8 moles of Mg and 4.7 moles of O_2?

$$2Mg + O_2 \rightarrow 2MgO$$

What would be the limiting reactant?

Which reactant is in excess, and how many moles of it are left over?

Practise

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3. What is the greatest amount of $AlCl_3$ (in grams) that can be made with 114 grams of Al and 186 grams of Cl_2?

$$2Al + 3Cl_2 \rightarrow 2\ AlCl_3$$

Inquiry Question 3:

How are chemicals in solutions measured?

2.3.1

- conduct practical investigations to determine the concentrations of solutions and investigate the different ways in which concentrations are measured (ACSCH046, ACSCH063)

Concentration

What are the different measurements of concentration?

- Mass of solute / litre of solution (________)
- Moles of solute / liter of solution (________)
- ppm (mg/kg or ________)
- ppb (µg/kg or __________)
- %w/w

  %v/v

  %w/v

Practise

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1. 1.32 g potassium permanganate, $KMnO_4$, was dissolved in water and the volume made accurately to 250mL. Calculate the concentration of this solution in g/100 mL and gL^{-1} of solution.

2. 13.6g copper sulfate was dissolved in 500g water. Calculate the %(w/w) copper sulfate in this solution.

3. What mass of sodium chloride has to be dissolved in 250 mL water to make a 0.90%(w/v) solution (the common saline solution in hospitals)? Assume that the solution has a volume of 250 mL.

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4. What volume of alcohol (ethanol) is present in 750mL of a 14%(v/v) solution of alcohol in water (for example a red wine)?

5. A solution contained 500ppm mercury(II) ion. Express this as a %(w/w)

6. 22.1 g potassium chloride was dissolved in water and the volume made accurately to 100mL. Calculate the concentration of the solution in grams per litre.

.

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7. 13.6 g sodium carbonate was dissolved in 250 g water. Calculate the %(w/w) of this solution.

8. What mass of iodine do you need to dissolve in 500mL water to make a 2.5%(w/w) solution? Assume the solution has the same volume as the volume of water used.

9. What volume of acetic acid is present in 500mL of a 6.0%(v/v) aqueous solution (a typical vinegar)?

10. A contaminated water supply contained 2.3ppm lead ion. What is the concentration of lead ions expressed as %(w/w)?

TO DO LIST:

- Textbook Questions Page 255

Inquiry Question 3:

How are chemicals in solutions measured?

2.3.2.a

- manipulate variables and solve problems to calculate concentration, mass or volume using:

  – $c = n\ V$ (molarity formula) (ACSCH063)

Molarity

What is molarity?

$c = n / v$

Molarity refers to the concentration of a solution in moles per litre.

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1. Calculate the molarity of a solution prepared by dissolving 9.8 moles of solid NaOH in enough water to make 3.62 L of solution.

2. How many moles of sodium hydroxide are in 38 mL of 0.50 mol/L NaOH?

3. How much 2.5 mol/L sulfuric acid should you use if you need 0.12 mol H_2SO_4?

4. You dissolve 152.5g of $CuCl_2$ in water to make a solution with a final volume of 2.25 L. What is its molarity?

5. A solution has a volume of 375 mL and contains 42.5 g of NaCl. What is its molarity?

Practise

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1. 17.54 g pure barium hydroxide was dissolved in water and made up to 500mL in a volumetric flask. Calculate the molarity of the solution.

2. What mass of pure sulfuric acid, H_2SO_4, must be dissolved in water and made up to 250mL in a volumetric flask to make a 0.550M solution?

3. Calculate the mass of ammonium sulfate needed to make 500 mL of a 0.304 mol L^{-1} solution.

4. A student made up 0.150 L of a 1.55 mol L-1 solution of magnesium chloride. Calculate the:
   a) number of moles of magnesium chloride in this solution.
   b) mass of solute in this solution.

Inquiry Question 3:

How are chemicals in solutions measured?

2.3.2.b

- manipulate variables and solve problems to calculate concentration, mass or volume using:

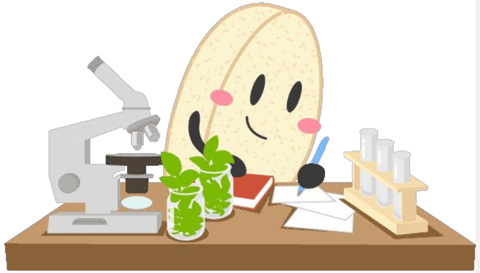
  – dilutions (number of moles before dilution = number of moles of sample after dilution)

Practical Demonstration

Dilutions

What stays the same and what changes?

Dilution Calculations



When a solution is diluted, the concentration is ___________________________, however the number of moles __________________________________.

$$C_1V_1 = C_2V_2$$

Example 1:

C

If 45.0 mL of a 6.00 M HCl solution are diluted to a final volume of 0.250 L, what is the final concentration?

Practise

makes perfect

1. How many mL of a 2.50 M NaOH solution are required to make 525 mL of a 0.150 M NaOH solution?

2. What volume of 12M H_2SO_4 will be needed to make 1.500 L of a 1.500 M solution of H_2SO_4?

3. To what volume should 50mL of a 1.50$molL^{-1}$ solution of potassium nitrate be diluted to make a solution that is 0.30$molL^{-1}$?

4. What is the concentration, in ppm, of a 0.00200M solution of HCl?

Inquiry Question 3:

How are chemicals in solutions measured?

2.3.3.

- conduct an investigation to make a standard solution and perform a dilution

Standard solutions

A standard solution is one whose **concentration is known accurately.**

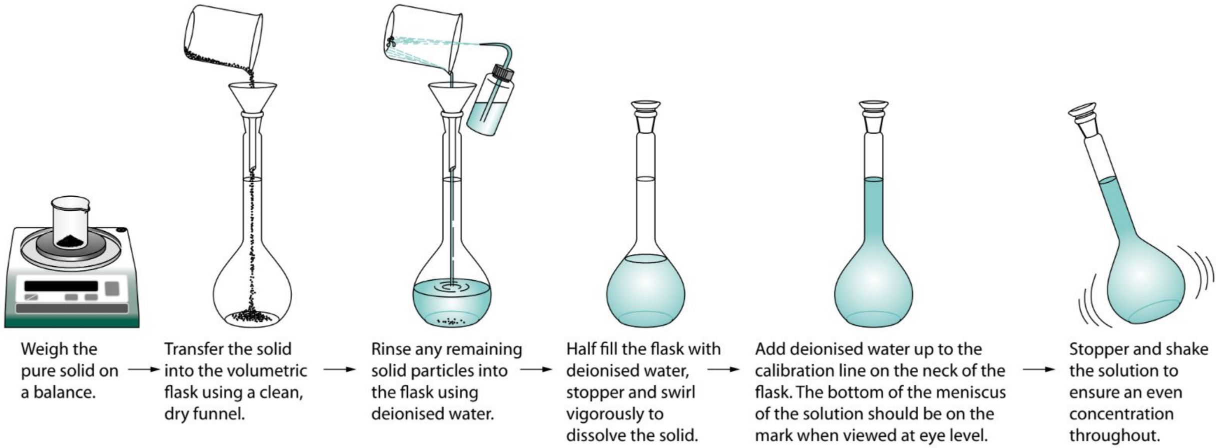
Primary

A **highly pure** solution made by dissolving an accurately measured mass of a solute in solvent.

Secondary

A solution that is **not pure** and specifically made for certain analysis.

Preparing a primary standard



Examples of primary standards:

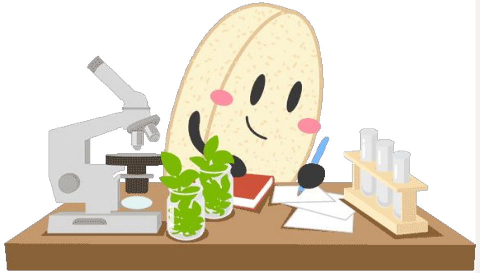
- **bases:** anhydrous sodium carbonate (Na2CO3) and hydrated sodium borate(Na2B4O7·10H2O)
- **acids:** hydrated oxalic acid (H2C2O4·2H2O) and potassium hydrogen phthalate (KH(C8H4O4)).

Practical Investigation

__Preparation of a Standard Solution__

Pearson 2.4

Concentrations of Standard Solutions



Example:

Calculate the concentration of a standard solution prepared from 117.0g of NaCl dissolved in a 500.0mL volumetric flask.

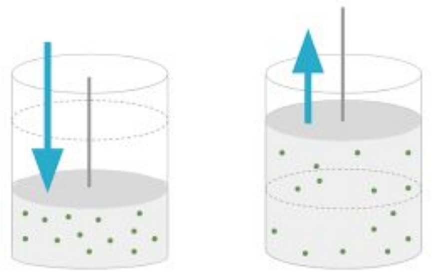
Inquiry Question 4:

How does the Ideal Gas Law relate to other Gas Laws?

2.4.1

- conduct investigations and solve problems to determine the relationship between the Ideal Gas Law and:
  - – Gay-Lussac's Law (temperature)
  - – Boyle's Law
  - – Charles' Law
  - – Avogadro's Law (ACSCH060)
-

Volume and Pressure



Volume is the quantity used to describe the space that a substance occupies.

Pressure is the force exerted on a surface by the particles of a gas as they collide with the surface.

Units of pressure:

Name of unit	Symbol for unit	Conversion to $N\,m^{-2}$
newtons per square metre	$N\,m^{-2}$	
pascal	Pa	$1\,Pa = 1\,N\,m^{-2}$
kilopascal	kPa	$1\,kPa = 1 \times 10^3\,Pa = 1 \times 10^3\,N\,m^{-2}$
atmosphere	atm	$1\,atm = 101.3\,kPa = 1.013 \times 10^5\,N\,m^{-2}$
bar	bar	$1\,bar = 100\,kPa = 1 \times 10^5\,N\,m^{-2}$
millimetres of mercury	mmHg	$760\,mmHg = 1\,atm = 1.013 \times 10^5\,N\,m^{-2}$

$1\,bar = 100\,kPa = 1 \times 10^5\,N\,m^{-2}$
$1\,atm = 101.3\,kPa = 1.013 \times 10^5\,N\,m^{-2}$

Practical Investigation

Investigating Gas Laws

Pearson 2.6

• Boyle's Law - $V_1P_1 = V_2P_2$ V_1 & V_2 = volume (L)
P_1 & P_2 = pressure (atm)

Found that volume & pressure of a fixed amount are inversely proportional when temperature is constant.

volume decreased
pressure increased

volume increased
pressure decreased

P_1 & P_2] in any units
V_1 & V_2] as long as it is the same
T_1 & T_2 = temp. (K)

Combined gas law

$$\frac{P1V1}{T1} = \frac{P2V2}{T2}$$

1°C = 273°K

• Charle's Law - $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ V_1 & V_2 = volume (L)
T_1 & T_2 = temperature (K)

Found that when pressure is constant, volume & temperature is proportional.

temperature increase
volume increase

↑T ↓T

temperature decrease
volume decrease

• Gay-Lussac's Law (pnet)

$\frac{P_1}{T_1} = \frac{P_2}{T_2}$ P_1 & P_2 = pressure (atm)
T_1 & T_2 = temperature (K)

Found that when volume is constant, temperature & pressure is proportional

temperature increase
pressure increase

↑T ↓T

temperature decrease
volume decrease

$\frac{V_1}{n_2} = \frac{V_2}{n_2}$

volume & moles proportional.

• Avogadro's Law → known as molar volume of gases

Showed that a mole of any gas will occupy same space under constant temperature & pressure

AT STP (standard temp & pressure) → 0°C & 100kPa
exactly 22.71L

1atm = 101.325 kPa

AT STL (standard lab conditions) → 25° & 100kPa
exactly 24.79L

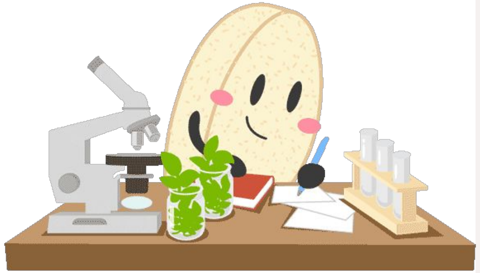
Ideal gas law

$$PV = nRT$$

→ use this to calculate the volume of gases consumed & produced in chemical reactions

P = pressure (kPa)
V = volume (L)
n = number of moles (mol)
R = universal gas constant (8.314J/mol/K)
T = temperature (K)

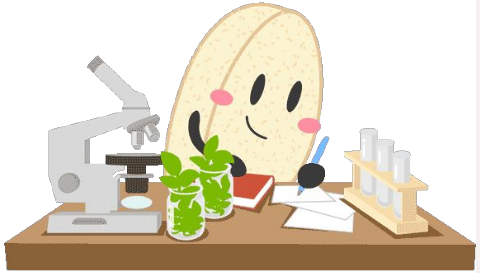
Mass - Volume Stoichiometry



Example:

Calculate the volume of carbon dioxide, in L, produced when 2.00kg of propane is burned completely in oxygen to produce carbon dioxide and water. The gas volume is measured at SLC.

Volume - Volume Stoichiometry



Example:

If 50mL of methane is burned, calculate the volume of O2 gas required for complete combustion of the methane under constant temperature and pressure conditions.