UNIT 3

AMOUNT OF SUBSTANCE AND MEASUREMENT

Student Version

$$6CO_2 + 6H_2O \xrightarrow[\text{nutrients}]{\text{sunlight}} \underbrace{C_6H_{12}O_6}_{\text{Organic compounds}} + 6O_2$$

Contents

Key words: base quantity, base unit, derived quantity, derived unit, thermometer, stop-clock, gas syringe, mass balance, stop-clock, measuring cylinder, pipette, burette, volumetric flask, density, mole, Avogadro's number, molarity, concentration, concentrated, dilute, standard solution, dilution factor, Avogadro's Law, ideal gas equation, molar gas constant, relative molecular mass, relative formula mass, empirical formula, reactant, product, stoichiometric coefficient, law of conservation of mass, laws of chemical combination

Units which must be completed before this unit can be attempted:
Unit 1 – Atoms and the Periodic Table
Unit 2 – Particles, Bonding and Structures

Estimated Teaching Time: 11 hours

The Periodic Table of the Elements

Key

relative atomic mass
symbol
name
atomic (proton) number

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

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

Lesson 1 – Why are practicals important?

a) Introduction

- Chemistry is a practical subject; practical skills are needed in order to:
  - Prepare, extract and purify substances (**synthesis**)
  - Identify substances (**qualitative analysis**)
  - Determine how much of a substance is present (**quantitative analysis**)

- When carrying out practical work in Chemistry, it is important to recognise that many chemicals are hazardous and can cause harm when they come into contact with skin, eyes or even clothes; **Safety precautions** are therefore very important when carrying out practical work in Chemistry

- Quantitative Analysis requires you to make measurements of various quantities; when making measurements, it is important to:
  - appreciate that most measurable quantities have **units**; understand what these units are and how the units can be interconverted
  - appreciate that measuring requires the use of specific **instruments**; understand what these instruments are and how they should be used
  - appreciate that no measuring instrument is perfect and that **errors** can arise both from the instruments themselves and how they are used

b) Safety Precautions in the Laboratory

- Here is a summary of the most important safety precautions and the reasons for them:

PRECAUTION	WHEN	REASON
Wear a lab coat	When you are carrying out practical work	To prevent your clothes from coming into contact with harmful chemicals
Avoid shorts or short skirts and avoid open footwear	Always	To prevent legs and feet from coming into contact with harmful chemicals
Wear safety goggles	Whenever you or anyone else is carrying out practical work	To protect your eyes from harmful chemicals
Wear disposable gloves	Whenever you or anyone else is dealing with chemicals known to be harmful to skin	To protect your hands form harmful chemicals
If you have long hair, wear it tied back	Always	To prevent hair from coming into contact with harmful chemicals or open flames
Don't eat or drink	Always	To prevent food contamination
Label all containers you are using to store chemicals	Whenever you are using chemicals	To ensure you don't get the chemicals mixed up
Follow all instructions and don't do anything you are not instructed to do	Always	To ensure you don't put yourself or anyone else in danger
Report all accidents, breakages and spillages to the teacher in charge immediately	As soon as you have an accident or break or spill someone	To allow the teacher to ensure that any injuries can be dealt with immediately, and that anything unsafe can be removed from the laboratory
Don't leave any fragile apparatus in a place where it might roll off or be knocked off the bench	Whenever you are using fragile apparatus, especially glassware	To avoid breaking fragile apparatus, which can be both dangerous and expensive
Always rinse all glassware, leave it to dry and wipe all surfaces after use	After any practical work	To ensure that the glassware and work surfaces contain no harmful chemical after use
Keep the lab free of clutter	Always	So that people have more space in which to work and move around safely
Never take any chemicals or equipment out of the laboratory	Always	They are not designed to be used outside the laboratory and are not safe to use anywhere else

Activity 1.1: Understanding risks and safety precautions in the laboratory

Design a poster demonstrating how to work safely in a laboratory. Include at least five important safety rules, covering at least one from the main categories – clothing, behaviour and communication. Include as many illustrations as possible!

c) Quantities, Units and Measuring Instruments

Summary Activity 1.2: units of temperature

- What are the different units of temperature commonly used?
- Express the following quantities in K: 25 ^oC, 100 ^oC, -273 ^oC
- Express the following quantities in ^oC: 345 K, 600 K, 100 K

(i) Important Quantities and Units in Chemistry

- In Chemistry, we measure lots of different quantities; the quantities we measure most frequently are:
  - volume
  - mass
  - temperature
  - time

- Each of these quantities has different units; there is an official international unit, called an SI unit, but there are also other units which are commonly used

Quantity (symbol)	SI unit	Other units commonly used in Chemistry
Volume (V)	cubic metre (m^3)	cubic centimetre (cm^3) (or ml) cubic decimetre (dm^3) (or L) 1000000 cm^3 = 1000 dm^3 = 1 m^3
Mass (m)	kilogram (kg)	gram (g) (1000 g = 1 kg)
Time (t)	second (s)	
Temperature (T)	Kelvin (K)	degrees celcius (^oC) T (K) = T (^oC) + 273

Test your knowledge 1.3: Interconverting important units in Chemistry

a) Express the following quantities in g:
   (i) 25 kg (ii) 3.2 kg (iii) 0.34 kg

b) Express the following quantities in m^3:
   (i) 25 cm^3 (ii) 3.2 dm^3 (iii) 0.34 dm^3 (iv) 150 cm^3 (v) 120 dm^3

c) Express the following quantities in dm^3:
   (i) 0.25 m^3 (ii) 3.2 m^3 (iii) 25 cm^3 (iv) 150 cm^3 (v) 6.2 cm^3

d) Express the following quantities in cm^3:
   (i) 0.25 m^3 (ii) 3.2 m^3 (iii) 0.40 dm^3 (iv) 0.015 dm^3 (v) 6.2 dm^3

Lesson 2 – What is a base quantity and what is a derived quantity?

(ii) Based and Derived Quantities and Units

- A **base quantity** is a quantity which cannot be calculated by combining other base quantities; there are seven base quantities in science, each of which has a **base unit:**

Base quantity (symbol)	Base SI unit	Used in Chemistry	Other units commonly used in Chemistry
Length (l)	metre (m)	No	
Mass (m)	kilogram (kg)	Yes	gram (g) (1000 g = 1 kg)
Time (t)	second (s)	Yes	
Temperature (T)	Kelvin (K)	Yes	degrees celcius (°C) T (K) = T (°C) + 273
Amount of substance (n)	mole (mol)	Yes	
Current (I)	Amp (A)	No	
Luminous Intensity	Candela (cd)	No	

- All other quantities are known as **derived quantities** - they depend on at least one of the base quantities and can be calculated if the base quantities are known; the units for derived quantities are called **derived units** and can be expressed in terms of the base units; the derived quantities which may need to be measured or used in Chemistry are:

Quantity (symbol)	Derived SI unit	Expressed in base units	Other units commonly used in Chemistry
Volume (V)	cubic metre (m^3)	m^3	cubic centimetre (cm^3) (or ml) cubic decimetre (dm^3) (or L) 1000000 cm^3 = 1000 dm^3 = 1 m^3
Pressure (P)	Pascal (Pa)	$kgm^{-1}s^{-2}$	atmosphere (atm) kilopascal (kPa) 1 kPa = 1000 Pa
Energy (E) (or H)	Joule (J)	kgm^2s^{-2}	kilojoule (kJ)
Voltage (V) (or pd) (or emf)	Volt (V)	$kgm^2s^{-3}A^{-1}$	
Charge (q)	Coulomb (C)	As	Faraday (F)

Test your knowledge 2.1: Using base and derived quantities

Express the following quantities in base units:

(a) Force (= mass x acceleration)
(b) Work done (= pressure x volume)
(c) Power (= voltage x current)
(d) Momentum (= mass x velocity)
(e) Rate of reaction (= concentration / time)

d) Measuring Instruments in Chemistry

(i) Different Measuring Instruments

- Special instruments are available for measuring quantities in Chemistry; each instrument has a specific purpose and comes with a given measurement error; the smaller the error, the more accurate the measurement that the instrument can make

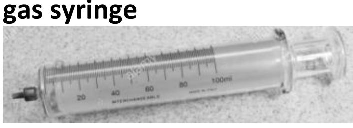
- One important quantity we need to be able to measure is mass:

Quantity	Instrument	Details
Mass (g)	mass balance 	Mass balances measure mass to 1 dp or 2 dp, depending on the balance. Typical error: ±0.1 g (if 1 dp), ±0.01 g (if 2 dp)

- Temperature and time are measured using the following instruments:

Quantity	Instrument	Details
Temperature (°C)	Thermometer 	Thermometers usually measure temperature to the nearest 0.5 °C Typical error: ±0.5 °C
Time (s)	stop-clock 	Stop-clocks can measure time to the nearest 0.01 s, but human reaction times are much longer than that, so human error is usually the limiting factor

- The quantity most frequently measured in Chemistry is **volume**; there are lots of different instruments available to measure volume; each instrument serves a slightly different purpose:

Instrument	Details	Typical error
measuring cylinder 	- measuring cylinders are convenient, but not accurate, ways of measuring volume - they should be used only when an approximate volume measurement is sufficient - they are most commonly found in sizes of 10 cm^3, 25 cm^3, 50 cm^3 and 100 cm^3	for 10 cm^3: ± 0.2 cm^3 for 25 cm^3: ± 0.5 cm^3 for 50 cm^3: ± 1 cm^3 for 100 cm^3: ± 2 cm^3
Pipette 	- most pipettes can only measure a single volume, usually 25 cm^3 - they can, however, measure this volume very accurately - they are designed to deliver a precise amount of liquid into another container - they should only be used with a **pipette filler**	for 25 cm^3: ± 0.05 cm^3
Burette 	- burettes are designed to deliver any volume up to 50 cm^3 - they are more accurate than measuring cylinders but less accurate than pipettes - the volume delivered can be deduced by subtracting the initial measurement from the final measurement - they need to be used with a stand, clamp and boss and they are mainly used in titrations	± 0.15 cm^3
volumetric flask 	- volumetric flasks are designed to measure a single, specific volume very accurately - they are not designed to deliver this volume; just to contain it - most volumetric flasks are designed to contain 250 cm^3 - they are mainly used to prepare standard solutions.	± 0.2 cm^3
gas syringe 	- gas syringes are used to collect and measure gas volumes - they have a similar accuracy to measuring cylinders	± 1 cm^3

Test your knowledge 2.2: Measuring Volumes

Copy and complete following table, stating the main advantage and disadvantage of each instrument for measuring volume:

Instrument	Advantage	Disadvantage
Pipette		
Volumetric flask		
Burette		
Measuring cylinder		

Lesson 3 – What is density and how can we measure it?

(ii) Measuring Densities

- The density of a substance is its mass per unit volume; the SI units of density are kgm^{-3}, but in the laboratory it is more common to measure density in gcm^{-3}
- DENSITY (gcm^{-3}) = MASS (g) /VOLUME (cm^3)

Practical 3.1: Comparing the densities of pure water and salt water

1) Weigh an empty 100 cm^3 measuring cylinder and record its mass.
2) Add 50 cm^3 of distilled water to the measuring cylinder.
3) Weigh the measuring cylinder again, this time with the water in it.
4) Determine the mass of the water in the measuring cylinder by subtracting the initial mass from the final mass.
5) Repeat steps (2) to (4) using salty water.
6) Copy and complete the following table:

	Pure water	Salty water
Mass of empty measuring cylinder (g)		
Mass of measuring cylinder with water (g)		
Mass of water (g)		
Volume of water (cm^3)		
Density of water (gcm^{-3})		

7) Hence calculate the density of the pure water and the salty water.

The density of pure water is 1.0 gcm^{-3}. How did this compare to your calculation?
Salty water is more dense than pure water. Is this what you discovered? Why do you think this is?

Practical 3.2: Measuring the density of sand

1) Weigh an empty 100 cm^3 measuring cylinder and record its mass.
2) Add sand to the measuring cylinder until it is approximately 20% full (ie the top of the sand is close to the 20 cm^3 mark).
3) Weigh the measuring cylinder again, this time with the sand in it.
4) Determine the mass of the sand in the measuring cylinder by subtracting the initial mass from the final mass.
5) Take another 100 cm^3 measuring cylinder and water until it is approximately half-full. Record the exact volume of water.
6) Pour the sand into the water and record the new level of the water.
7) Determine the volume of the sand in the measuring cylinder by subtracting the initial mass from the final mass.
8) Copy and complete the following table:

Mass of empty measuring cylinder (g)	
Mass of measuring cylinder with sand (g)	
Mass of sand (g)	
Volume of water without sand (cm^3)	
Volume of water with sand (cm^3)	
Volume of sand (cm^3)	
Density of sand (gcm^{-3})	

9) Hence calculate the density of sand.

The density of sand varies between 1.4 gcm^{-3} and 1.6 gcm^{-3} depending on the type of sand. Is your measurement within these limits?
How does the density of sand compare with water? Could you have predicted this?
What do you think are the main sources of error in this experiment?

e) The Mole

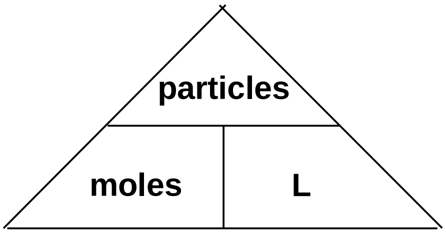
(i) Avogadro's number

- Atoms and molecules are very small – far too small to count individually! Since they are so small, any sensible laboratory quantity of substance must contain a huge number of individual particles:
  - 1 litre of water contains 3.3 x 10^{25} molecules
  - 1 gram of magnesium contains 2.5 x 10^{22} atoms
  - 100 cm^3 of oxygen contains 2.5 x 10^{21} molecules

- Such numbers are not convenient to work with, but it very important in Chemistry to be able to measure the amount of substance present; it is therefore necessary to find a unit of "amount" which corresponds better to the sort of quantities of substance normally being measured; the unit chosen for this purpose is the **mole**
- The number of particles in one mole of a substance is 6.02 x 10^{23}; this is known as **Avogadro's number, L**; when we need to know the number of particles of a substance, we usually count the number of moles; it is much easier than counting the number of particles

- The number of particles can be calculated by multiplying the number of moles by Avogadro's number; the number of moles can be calculated by dividing the number of particles by Avogadro's number

(Number of particles) = (number of moles) x L



Test your knowledge 3.3: Using Avogadro's number

a) If you have 2.5 x 10^{21} atoms of magnesium, how many moles do you have?
b) If you have 0.25 moles of carbon dioxide, how many molecules do you have?
c) How many moles are present in 3 x 10^{22} molecules of nitrogen?
d) How many atoms of carbon are present in 0.02 moles?
e) If you have 9.0 x 10^{24} molecules of oxygen, how many moles do you have?

Lesson 4 – How can we work out how many moles we have in a sample?

(ii) Molecular and formula masses

- Elements and compounds all have either a **molecular formula** (if they have a simple molecular or simple atomic structure) or a **unit formula** (if they have a giant structure)

- Atoms have a **relative atomic mass**, which is given in the Periodic Table

- Simple molecular substances have a relative **molecular mass**, which is the sum of the relative atomic masses of all the atoms in one molecule of that substance; it can also be defined in the same way as relative atomic mass:

The relative molecular mass of a molecule is the ratio of the average mass of that molecule to 1/12th of the mass of one atom of carbon-12

- Giant structures have a **relative formula mass**, which is the sum of the relative atomic masses of all the atoms in one formula unit of that substance

The relative formula mass of a giant structure is the ratio of the average mass of one formula unit of that substance to 1/12th of the mass of one atom of carbon-12

Test your knowledge 4.1: Deducing relative molecular masses and relative formula masses

(a) Use the Periodic Table of Elements to write down the atomic masses of:
  (i) Carbon (ii) Oxygen (iii) Chlorine (iv) Sodium (v) Hydrogen (vi) Magnesium
(b) Use your answers from question 1 to deduce the molecular masses of:
  (i) O_2 (ii) CO_2 (iii) Cl_2 (iv) HCl (v) CH_4 (vi) H_2O
(c) Use your answers from question 1 to deduce the formula masses of:
  (i) NaCl (ii) Na_2CO_3 (iii) MgO (iv) $MgCl_2$ (v) $Mg(OH)_2$

(iii) Moles and molar masses

- The number of particles in a mole (Avogadro's number) is chosen so that 1 mole of a substance corresponds to its relative mass measured in grams
  - one mole of carbon has a mass of 12.0 g
  - one mole of hydrogen atoms has a mass of 1.0 g
  - one mole of hydrogen molecules has a mass of 2.0 g
  - one mole of sodium chloride has a mass of 58.5 g

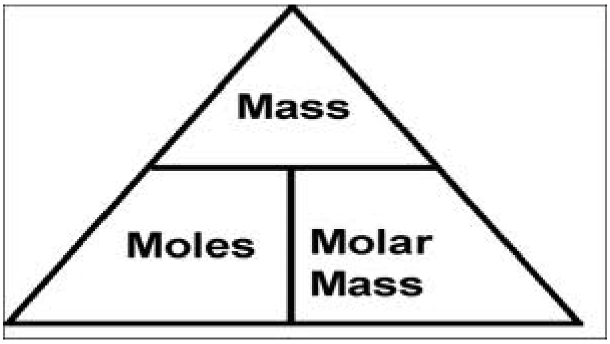
- This is how the mole is defined: "a mole of a substance is the amount of that substance which contains the same number of particles as there are in 12.0 grams of carbon-12"

- The mass of one mole of a substance is known as its **molar mass** and has units of $gmol^{-1}$; the symbol for the molar mass is m_r

- You can find the number of moles of a substance if you are given its **mass** and you know its **molar mass**:

number of moles = mass/molar mass

$$n = m/m_r$$



Example: Solution:	Calculate the number of moles of carbon present in 3 g of carbon moles = mass/molar mass = 3/12 = 0.25
Example: Solution:	Calculate the mass of 0.2 moles of NaOH molar mass = 23 + 16 + 1 = 40; mass = moles x molar mass = 0.2 x 40 = 8 g
Example: Solution:	Calculate the molar mass of a substance if 0.1 moles of that substance has a mass of 3.2 g molar mass = mass / moles = 3.2/0.1= 32 $gmol^{-1}$

Test your knowledge 4.2: Using mass measurements to calculate moles

a) Calculate the number of moles present in:	b) Calculate the mass of:	c) Calculate the molar mass of the following substances:
(i) 2.3 g of Na	(i) 0.05 moles of Cl_2	(i) 0.015 moles, 0.42 g
(ii) 2.5 g of O_2	(ii) 0.125 moles of KBr	(ii) 0.0125 moles, 0.50 g
(iii) 240 kg of CO_2	(iii) 0.075 moles of $Ca(OH)_2$	(iii) 0.55 moles, 88 g
(iv) 12.5 g of $Al(OH)_3$	(iv) 250 moles of Fe_2O_3	(iv) 2.25 moles, 63 g
(v) 5.2 g of PbO_2	(v) 0.02 moles of $Al_2(SO_4)_3$	0.00125 moles, 0.312 g

Lesson 5 – How can we work out how many moles we have in a solution?

(iv) Amount of Substance in Solution

- The amount of substance present in a given volume of solution is known as the **concentration** of the solution; concentration is usually measured in moles per cubic decimetre ($moldm^{-3}$); this specific type of concentration measurement is also known as the **molar concentration** or **molarity** of the solution

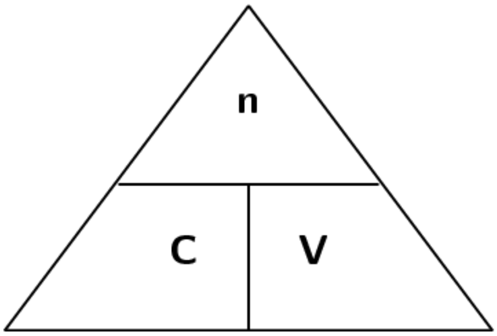
- Concentration can also be measured in grams per cubic decimetre (gdm^{-3}); this type of concentration measurement is also known as the **mass concentration** of the solution

- A solution with a relatively high concentration is said to be concentrated; a solution with a relatively low concentration is said to be dilute
  - a typical solution of hydrochloric acid used in the laboratory will have a molarity of 0.1 $moldm^{-3}$; this would be considered to be dilute hydrochloric acid
  - the maximum possible concentration of hydrochloric acid has a molarity of 12 $moldm^{-3}$; this is known as concentrated hydrochloric acid

- You can find the number of moles of a substance dissolved in water (aqueous) if you are given the **volume** of solution and you know its **molar concentration**:

number of moles = aqueous volume x molar concentration

n = V x C



Aqueous volume MUST be measured in dm^3
concentration has units of $moldm^{-3}$

Example: Solution:	Calculate the number of moles present in 20 cm^3 of a 0.1 $moldm^{-3}$ solution moles = molarity x volume (dm^3) = 0.1 x 20/1000 = 0.002 moles molar mass = 23 + 16 + 1 = 40 $gmol^{-1}$, so mass = moles x molar mass = 40 x 0.125 = 5 g
Example: Solution:	Calculate the molarity of a solution containing 0.01 moles of solute in 50 cm^3 of solution molarity = moles / volume (dm^3) = 0.01 / (50/1000) = 0.2 $moldm^{-3}$
Example: Solution:	Calculate the mass of NaOH required to make 250 cm^3 of a 0.5 $moldm^{-3}$ solution moles required = molarity x volume (dm^3) = 0.5 x 250/1000 = 0.125 moles molar mass = 23 + 16 + 1 = 40 $gmol^{-1}$, so mass = moles x molar mass = 40 x 0.125 = 5 g

Test your knowledge 5.1: Using moles, molarity and aqueous volume

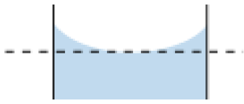
a) Calculate the number of moles of substance present in each of the following solutions:	b) Calculate the molarity of the following solutions:	c) Calculate the molarity of the following solutions:
(i) 25 cm^3 of 0.1 $moldm^{-3}$ HCl	(i) 0.05 moles of HCl in 20 cm^3	(i) 35 g of NaCl in 100 cm^3
(ii) 40 cm^3 of 0.2 $moldm^{-3}$ HNO_3	(ii) 0.01 moles of NaOH in 25 cm^3	(ii) 20 g of $CuSO_4$ in 200 cm^3
(iii) 10 cm^3 of 1.5 $moldm^{-3}$ NaCl	(iii) 0.002 moles of H_2SO_4 in 16.5 cm^3	(iii) 5 g of HCl in 50 cm^3
(iv) 5 cm^3 of 0.5 $moldm^{-3}$ $AgNO_3$	(iv) 0.02 moles of $CuSO_4$ in 200 cm^3	(iv) 8 g of NaOH in 250 cm^3
(v) 50 cm^3 of 0.1 $moldm^{-3}$ H_2SO_4	(v) 0.1 moles of NH_3 in 50 cm^3	(v) 2.5 g of NH_3 in 50 cm^3

(v) Standard Solutions

- A standard solution is a solution of precisely known concentration; standard solutions can be prepared by dissolving a measured mass of solute to make a specific volume of solution; standard solutions are important in chemical analysis

- A standard solution can be prepared using the following steps:
  - weigh out the required mass of solute (weigh an empty weighing boat, then add the correct mass of solid to the weighing boat using a spatula)
  - pour the solid into a beaker
  - add enough water to dissolve the solid (usually no more than 50 cm^3 of water; use a stirring rod to help dissolve the solute and mix the contents well)
  - transfer the solution to a volumetric flask using a funnel
  - make up to the mark with distilled water (rinse out the beaker a few times with distilled water and add the washings to the volumetric flask)

Note: when the level of water reaches the neck of the flask, you will notice that the surface of the water is not completely flat, but slightly curved:



This curvature is called the meniscus; for an accurate measurement of volume, the base of the meniscus should lie completely on the line, as in the diagram above; this is true of all measuring devices

- Based on the volume and molarity of standard solution to be prepared, you can calculate the mass of solute to be added as follows:

  mass = moles x molar mass and moles = molarity x volume

  so mass = molar mass x molarity x volume

  Eg If you want to prepare 250 cm^3 of 0.2 $moldm^{-3}$ NaOH, you will need 40 x 0.2 x 250/1000 x 0.2 = 2.0 g

  so **the required mass of NaOH would be 2.0 g**

Practical 5.2: Preparing 250 cm^3 of 0.1 $moldm^{-3}$ standard solutions of sodium chloride (NaCl) and sugar ($C_{12}H_{22}O_{11}$)

1) Deduce the mass of each solid required to prepare 250 cm^3 of a 0.1 $moldm^{-3}$ solution
2) Weigh out the required amount of each solid using a weighing boat, and pour it into a beaker
3) Add enough water to completely dissolve the solid, and then transfer the solution into a volumetric flask
4) Add distilled water, including washings from the beaker, until the base of the meniscus lies on the graduated mark on the volumetric flask, shaking well

Lesson 6 – How can we prepare standard solutions by diluting concentrated solutions?

- If the percentage by mass of solute in the concentrated solution is known, it is possible to calculate the mass of concentrated solution needed to make a certain volume of a more dilute solution:
  - Deduce the number of moles of solute required in the final solution (from the desired molarity and volume)
  - Deduce the mass of solute required to provide this number of moles
  - Use the percentage by mass of solute in the concentrated solution to deduce the mass of concentrated solution required to provided this mass of solute:

$$\text{Required mass of concentrated solution} = \frac{\text{required mass of solute x 100}}{\text{\% by mass of solute in solution}}$$

Example:	Concentrated HCl is known to contain 36% HCl by mass. What mass of concentrated HCl is required to make 250 cm^3 of a 0.1 $moldm^{-3}$ solution?
Solution:	The moles of required is 250/1000 x 0.1 = 0.025 so the mass of HCl required is 0.025 x 36.5 = 0.9125 g so the mass of concentrated HCl which contains this mass of pure HCl = 0.9125 x 100/36 = 2.53 g so the required mass of concentrated HCl needed would be **2.53 g**

Demonstration 6.1: Preparing 250 cm^3 of a 0.1 moldm^{-3} standard solution of HCl from a sample of concentrated HCl

(CAUTION – concentrated HCl is highly corrosive)

1) Weigh out 2.53 g of concentrated HCl using a weighing bottle
2) Add 100 cm^3 of water to a beaker, and then add the HCl and stir (never add water to concentrated acid – always add acid to water)
3) Transfer the solution into a volumetric flask and add water until the base of the meniscus lies on the graduated mark on the volumetric flask, shaking well and including washings from the weighing bottle and the beaker

- If the percentage by mass of solute is not known accurately, it is necessary to measure a volume of concentrated solution instead; the correct volume of solution to be diluted can be calculated using the dilution factor:

$$\textbf{dilution factor} = \frac{\text{molarity of original (concentrated) solution}}{\text{molarity of final (diluted) solution}}$$

$$\text{The volume of original concentrated solution to be diluted} = \frac{\text{volume of final (diluted) solution}}{\text{dilution factor}}$$

Worked example:	A 2.0 moldm^{-3} solution of hydrogen peroxide (H_2O_2) needs to be diluted to make 250 cm^3 of a 0.05 moldm^{-3} solution. What volume of the original 2.0 moldm^{-3} solution should be diluted?
Solution:	The dilution factor is 2.0/0.05 = 40 The desired volume of 0.05 moldm^{-3} is H_2O_2 is 250 cm^3 so the volume of 2.0 moldm^{-3} solution of H_2O_2 needed is 250/40 = 6.25 cm^3 **So 6.25 cm^3 of H_2O_2 should be diluted to a volume of 250 cm^3 in order to achieve this dilution**
Alternative solution:	number of moles required = C x V = 250/1000 x 0.05 = 0.0125 volume of the original solution containing this number of moles = n/C = 0.0125/2 = 0.00625 dm^3 = 6.25 cm^3 **So 6.25 cm^3 of H_2O_2 should be diluted to a volume of 250 cm^3 in order to achieve this dilution**

Practical 6.2: Prepare 250 cm^3 of a 0.1 moldm^{-3} solution of hydrogen peroxide by diluting a 2.0 moldm^{-3} solution

1) Work out the volume of 2.0 moldm^{-3} H_2O_2 required for the dilution
2) Measure this volume of 2.0 moldm^{-3} H_2O_2 as accurately as possible into a measuring cylinder
3) Add 100 cm^3 of water to a beaker, and then add the H_2O_2 and stir
4) Transfer the solution into a 250 cm^3 volumetric flask and add water until the base of the meniscus lies on the graduated mark on the volumetric flask, shaking well

Test your knowledge 6.3: Preparing standard solutions by dilution

(a) Concentrated nitric acid contains 65% HNO_3 by mass. What mass of concentrated nitric acid should be diluted to make 250 cm of a 0.1 moldm^{3-} solution?

(b) What volume of water should be added to 5.0 cm^3 of 6.0 moldm^{-3} NaOH to make a solution with molarity 0.10 moldm^{-3}?

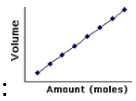
Lesson 7 – How can calculate the moles present in a gaseous sample?

(vi) Avogadro's Law and the Ideal Gas Equation

Summary Activity 7.1: The Gas Laws

- Why do all gases have similar physical properties?
- What is meant by the term "atmospheric pressure" and what is its value?
- What is the combined gas law?
- Which laws are used to create the combined gas law?

- The combined gas law states that $P_1V_1/T_1 = P_2V_2/T_2$; this means that for a fixed amount of any gas, PV/T = k, at least approximately; so a fixed amount of any gas at the same temperature and pressure will have the same volume

- The greater the number of moles of particles, the greater the volume that the gas will occupy at a given temperature and pressure; the volume occupied by a gas is directly proportional to the number of moles at a given temperature and pressure; this is known as **Avogadro's Law**:
  - Mathematically: $\frac{V}{n} = k$ or $\frac{\mathrm{V1}}{\mathrm{n1}} = \frac{\mathrm{V2}}{\mathrm{n2}}$
  - Graphically: 

Example:	If 0.02 moles of a gas occupy a volume of 0.4 dm^3, what volume will 0.03 moles of gas occupy at the same temperature and pressure?
Answer:	$V_1/n_1 = 0.4/0.02 = 20 = V_2/n_2$, so $V_2 = 20 \times 0.03 =$ **0.6 dm^3**

Test your knowledge 7.2: Using Avogadro's Law

0.1 moles of oxygen are found to occupy a volume of 2.4 dm^3 at room temperature and pressure.

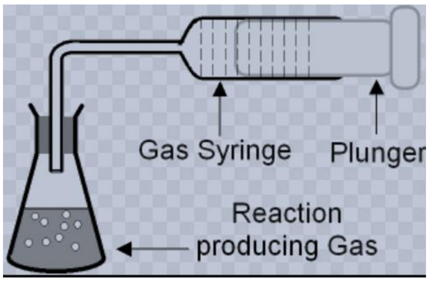
(a) What volume will be occupied by 0.1 moles of carbon dioxide under the same conditions?
(b) What volume will be occupied by 0.3 moles of nitrogen under the same conditions?
(c) What volume will be occupied by 1 mole of chlorine under the same conditions?
(d) How many moles of argon will be required to fill a container of volume 12 dm^3 under the same conditions?
(e) How many moles of hydrogen are present in 120 cm^3 under the same conditions?

- Avogadro's Law can be combined with the combined gas law to produce the following: PV/nT = k
  - If P is the pressure measured in pascals (Pa), V is the volume in m^3, T is the temperature measured in Kelvin (K), and n is the number of moles, then the value of the constant is 8.31 $Jmol^{-1}K^{-1}$
  - the constant is given the symbol R and is known as the **molar gas constant**
  - the pressure, temperature, volume and amount of gas can be related by a simple equation known as the **ideal gas equation**: **PV = nRT**
  - Note: volume must be measured in m^3: 1 m^3 = 1000 dm^3 = 10^6 cm^3
  - Note: temperature must be measured in Kelvin (K): 0 °C = 273 K

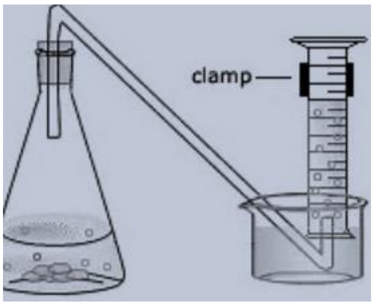
Term	Meaning	Units
P	Pressure	Pa
V	Volume	m^3
N	amount of substance	mol
T	Temperature	K

- Usually, experiments are carried out under normal pressure conditions; atmospheric pressure is approximately 100 kPa; this quantity is sometimes referred to as **one atmosphere**

- If the gas can be collected, its volume at atmospheric pressure (100 kPa) can be measured either using a gas syringe or using a measuring cylinder inverted over water:



www.gcsescience.com
using a gas syringe



www.socratic.org
using a measuring cylinder

Demonstration 7.3: Measuring the volume of a gas

Your teacher will carry out a simple reaction which produces a gas using one of the methods above.

Note how the volume of gas is measured.
Use the ideal gas equation to deduce the number of moles of gas evolved.

- Using this equation, it is possible to calculate the volume occupied by one mole of a gas at standard temperature (273 K) and pressure (1 atm = 101 kPa) (**stp** = standard temperature and pressure):
  - V = nRT/P = 1 x 8.31 x 273/101,000 = **0.0224 m^3 = 22.4 dm^3 = 22,400 cm^3**
  - This volume is known as the **molar volume of a gas at stp**

- It is also possible to calculate the volume occupied by one mole of a gas at room temperature (298 K) and pressure (1 atm = 101 kPa) (**rtp** = room temperature and pressure):

  - V = nRT/P = 1 x 8.31 x 298/101,000 = **0.0244 m^3 = 24.4 dm^3 = 24,400 cm^3**
  - This volume is known as the **molar volume of a gas at rtp**

Test your knowledge 7.4: Using the ideal gas equation

a) Calculate the number of moles present in:	b) Calculate the volume of gas occupied by:	c) Calculate the mass of the following gas samples:
(i) 48 dm^3 of O_2 at 298 K and 100 kPa	(i) 0.05 moles of Cl_2 at 298 K and 100 kPa	(i) 48 dm^3 of O_2 at 298 K and 100 kPa
(ii) 1.2 dm^3 of CO_2 at 298 K and 100 kPa	(ii) 0.25 moles of CO_2 at 298 K and 100 kPa	(ii) 1.2 dm^3 of CO_2 at 298 K and 100 kPa
(iii) 200 cm^3 of N_2 at 273 K and 250 kPa	(iii) 28 g of N_2 at 273 K and 250 kPa	(iii) 200 cm^3 of N_2 at 273 K and 250 kPa
(iv) 100 dm^3 of Cl_2 at 30 °C at 100 kPa	(iv) 3.2 g of O_2 at 30 °C at 100 kPa	(iv) 100 dm^3 of Cl_2 at 30 °C at 100 kPa
(v) 60 cm^3 of NO_2 at 25 °C and 100 kPa	(v) 20 g of NO_2 at 25 °C and 100 kPa	(v) 60 cm^3 of NO_2 at 25 °C and 100 kPa

Lesson 8 – What is an empirical formula and how is it different from a molecular formula or a unit formula?

f) Empirical Formulae

Summary Activity 8.1: Unit formula and molecular formula

- What is meant by the terms "molecular formula" and "unit formula"? Give one example of each.

(i) Definition of empirical formula

- The **empirical formula** of a compound is the formula which shows the **simplest whole-number ratio in which the atoms of each element in that compound exist**

- The empirical formula is not the same as the **unit formula** of a compound, which is the simplest whole-number ratio in which the particles in that compound exist; in giant covalent substances, the empirical formula and unit formula are always the same, but in ionic compounds they can be different:

Name of ionic compound	Unit formula	Empirical Formula
Sodium oxide	Na_2O (2 Na^+ ions per O^{2-} ion)	Na_2O
Sodium peroxide	Na_2O_2 (2 Na^+ ions per O_2^{2-} ion)	NaO
Aluminium hydroxide	$Al(OH)_3$ (3 OH^- ions per Al^{3+} ion)	AlO_3H_3
Ammonium Nitrate	NH_4NO_3 1 NH_4^+ ion per NO_3^- ion	$N_2H_4O_3$
Magnesium nitrate	$Mg(NO_3)_2$ 2 NO_3^- ions per Mg^{2+} ion	MgN_2O_6

- The empirical formula is also not the same as the **molecular formula** of a substance, which is the actual number of atoms of each element in one molecule of that substance:

Name of molecule	Molecular formula	Empirical Formula
Chlorine	Cl_2	Cl
Carbon dioxide	CO_2	CO_2
Ethane	C_2H_6	CH_3
Ethene	C_2H_4	CH_2
Propene	C_3H_6	CH_2

- The empirical formula does not uniquely identify a substance, because different substances can have the same empirical formula (eg C_2H_4 and C_3H_6, or NO_2 and N_2O_4)

(ii) Determining empirical formulae from mass composition data

- If the percentage composition by mass of each element in a compound is known, or the actual mass of each element in a known mass of compound is known, then its empirical formula can be determined by:
- dividing the mass of each element by its molar mass
- dividing each answer by the smallest of the answers
- multiplying all answers by the smallest factor required to ensure that all numbers are whole numbers

Example:	If a compound contains 85.8% carbon and 14.2% hydrogen, what is its empirical formula?
Answer:	Mole ratio = 85.8/12: 14.2/1 = 7.15:14.2 = 1:2 so **empirical formula = CH_2**

- The molecular formula is always a simple multiple of the empirical formula and can be therefore be deduced if the empirical formula and the relative molecular mass are known

Example:	If a compound has the empirical formula CH_2 and a relative molecular mass of 56, what is its molecular formula?
Answer:	relative empirical formula mass = 12 + 2 = 14, so there must be 56/14 = 4 empirical formula units in the molecular formula, so the **molecular formula** = $(CH_2)_4$ **= C_4H_8**

Test Your Knowledge 8.2: Empirical Formulae

a) A compound contains C 62.08%, H 10.34% and O 27.58% by mass. Find its empirical formula and its molecular formula given that its relative molecular mass is 58.
b) Find the empirical formula of the compound containing C 22.02%, H 4.59% and Br 73.39% by mass.
c) A compound containing 84.21% carbon and 15.79% hydrogen by mass has a relative molecular mass of 114. Find its molecular formula.
d) Analysis of a hydrocarbon showed that 7.8 g of the hydrocarbon contained 0.6 g of hydrogen and that the relative molecular mass was 78. Find the molecular formula of the hydrocarbon.
e) 3.36 g of iron join with 1.44 g of oxygen in an oxide of iron. What is the empirical formula of the oxide?
f) An ionic compound is analysed and found to contain 48.4% oxygen, 24.3% sulphur, 21.2% nitrogen and 6.1% hydrogen. Calculate its empirical formula and deduce its unit formula.

g) Equations

(i) Introducing Chemical Equations

- During chemical reactions, atoms and ions rearrange themselves and combine with other atoms and ions to form new substances; the starting substances in a chemical reaction are called the **reactants**, and the new substances created in a chemical reaction are called **products**

- The chemical changes taking place in a chemical reaction are best shown in a chemical equation; chemical equations always show:
  - the chemical formulae of the reactants and products
  - the ratio of the number of moles of reactants reacting together, and the number of moles of products made; these relative numbers are written in front of their respective formulae and are known as **stoichiometric coefficients**

- Chemical equations can also often show the state symbols of the reactants and products

- Consider the equation: $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$
  - In this reaction, carbon dioxide (CO_2) reacts with water (H_2O) to make glucose ($C_6H_{12}O_6$) and oxygen (O_2)
  - carbon dioxide and water are the reactants; glucose and water are the products
  - the stoichiometric coefficients are 6, 6, 1 and 6; they show that per mole of glucose made, six moles of oxygen are made and six moles of both carbon dioxide and water are required

- Note that stoichiometric coefficients do not show the actual amount of a substance which is reacting; they only show the ratio of the amounts of substance reacting; if you know the number of moles of any one substance involved in the reaction, use can use the chemical equation to deduce the number of moles of all of the other substances involved:

Example: Answer:	How many moles of water are needed to react with 0.03 moles of carbon dioxide? 6 moles of water react with 6 moles of carbon dioxide (1:1 ratio), so 0.03 moles of water are needed to react with 0.03 moles of carbon dioxide
Example: Answer:	How many moles of glucose can you make from 0.03 moles of carbon dioxide? 6 moles of carbon dioxide make 1 mole of glucose (6:1 ratio), so 0.03 moles of carbon dioxide will make 0.03/6 = 0.005 moles of glucose
Example: Answer:	How many moles of oxygen can you make from 0.03 moles of carbon dioxide? 6 moles of carbon dioxide make 6 moles of oxygen (1:1 ratio), so 0.03 moles of carbon dioxide will make 0.03 moles of oxygen

Test your knowledge 9.1: Using equations to calculate numbers of moles

a) Using the equation $Mg + 2HCl \rightarrow MgCl_2 + H_2$:
   (i) How many moles of magnesium would be needed to react with 0.01 moles of hydrochloric acid?
   (ii) How many moles of hydrogen could be produced from 0.01 moles of hydrochloric acid?
b) Using the equation $2\ H_2S + 3O_2$ ⍰ $2SO_2 + 2H_2O$:
   (i) How many moles of oxygen are needed to react with 0.5 moles of hydrogen sulphide?
   (ii) How many moles of sulphur dioxide can be made from 0.5 moles of hydrogen sulphide?
   c) Using the equation $4K + O_2$ ⍰ $2K_2O$:
   (i) How many moles of oxygen are needed to react with 0.05 moles of potassium?
   (ii) How many moles of potassium oxide can be made from 0.05 moles of potassium?

(ii) Law of Conservation of Mass

- During chemical reactions, particles are neither gained or lost; they are simply rearranged; if the reaction is taking place in a closed system, therefore, the total mass will remain constant over time, and the total mass of products will be the same as the total mass of reactants; this is known as the **Law of Conservation of Mass**

Online task 9.2: Illustrating the law of conservation of mass
If you have access to the internet, you can watch some illustrations of the law of conservation of mass:

(i) During the precipitation of barium sulphate or silver nitrate
www.youtube.com/watch?v=mcnga-bbNXk

(ii) During the reaction between vinegar and baking powder
www.youtube.com/watch?v=FZwHH7Sm4hI

- Together with the Law of Constant Composition and the Law of Multiple Proportions, these three laws are known as the **Laws of Chemical Combination**

(iii) Balancing Chemical Equations

- According to the Law of Conservation of Mass, the total number of each type of atom in the reactants should be equal to the total number of each type of atom in the products
For example, in this equation: $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$ there are 6 carbon atoms, 12 hydrogen atoms and 18 oxygen atoms on both sides of the equation

- If you know the chemical formulae of all of the reactants and products in the equation, you can use the Law of Conservation of Mass to deduce what the stoichiometric coefficients must be; this is known as **balancing an equation**

- When balancing an equation, balance compounds first, then elements

Example: Write a balanced chemical equation to show how magnesium (Mg) reacts with oxygen (O_2) to make magnesium oxide (MgO)

Answer: Mg and O_2 are the reactants; there are at least 2 O atoms on the left-hand side, so there must be at least 2 O atoms on the right-hand side, so the stoichiometric coefficient in front of MgO must be at least 2
this means there are at least 2 Mg atoms on the right-hand side, so there must also be at least 2 Mg atoms on the left-hand side, so the stoichiometric coefficient in front of Mg must also be at least 2
$2Mg + O_2 \rightarrow 2MgO$

Test your knowledge 9.3: Balancing Chemical Equations
The following chemical equations are not balanced. Balance them by adding the correct stoichiometric coefficients in front of the chemical formulae:

a) $N_2 + H_2 \rightarrow NH_3$
b) $Na + O_2 \rightarrow Na_2O$
c) $Al + Cl_2 \rightarrow AlCl_3$
d) $CH_4 + O_2 \rightarrow CO_2 + H_2O$
e) $HCl + O_2 \rightarrow Cl_2 + H_2O$

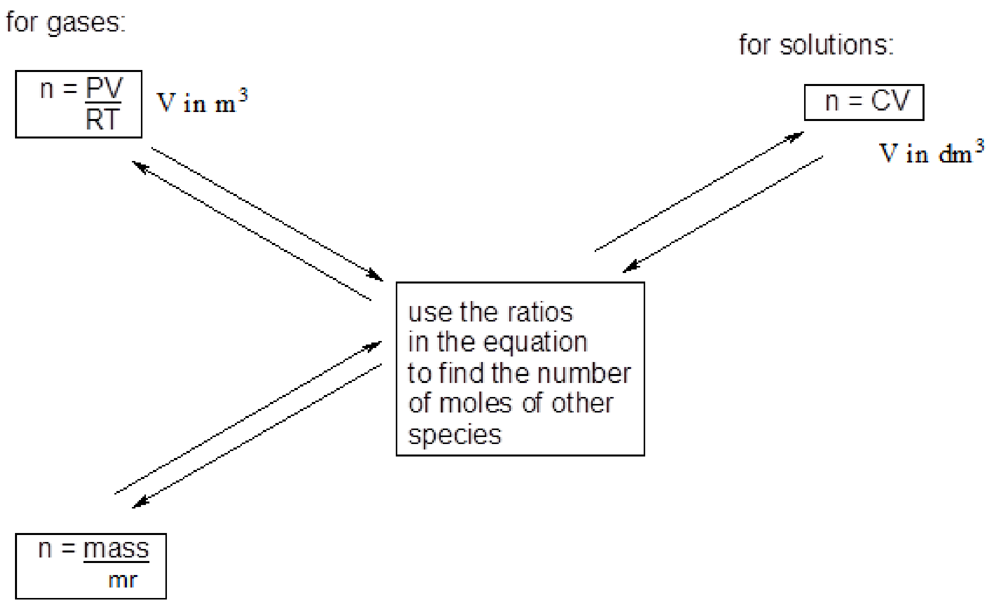
h) Calculating Reacting Quantities

- You can determine how many moles of a substance if you are given:
  - the mass of one substance in a chemical reaction, or
  - the volume and concentration of one aqueous substance in a chemical reaction, or
  - the volume, pressure and temperature of one gaseous substance in a chemical reaction

- If you have the number of moles of one substance in the reaction, you can use the stoichiometric coefficients to calculate the number of moles of everything else involved in the reaction; using these moles, you can calculate any other quantity you need

Example:	What volume (in dm^3) of hydrogen is produced at 298 K and 100 kPa when 192 g of magnesium is reacted with hydrochloric acid? $Mg + 2HCl \rightarrow MgCl_2 + H_2$
Answer:	moles of Mg = 192/24 = 8 Mg and H_2 react in a 1:1 ratio so moles of H_2 = 8 So volume of H_2 = 8 x 8.31 x 298/100000 = 0.198 m^3 = **198 dm^3**
Example:	What volume (in cm^3) of 0.2 $moldm^{-3}$ HCl is needed to react completely with 5.0 g of $CaCO_3$? $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$
Answer:	moles of $CaCO_3$ = 5/100 = 0.05 $CaCO_3$ and HCl react in a 1:2 ratio so moles of HCl = 0.05 x 2 = 0.1 So volume of HCl = n/C = 0.1/2 = 0.05 dm^3 = **50 cm^3**

- The relationships between the important quantities are shown below:



Test your knowledge 10.1: Calculating Reacting Quantities

a) What volume (in cm^3) of 0.5 $moldm^{-3}$ hydrochloric acid is required to react completely with 1.94 g of magnesium? $Mg + 2HCl \rightarrow MgCl_2 + H_2$

b) What volume (in dm^3) of oxygen at 298 K and 100 kPa is needed to react with 8.5 g of hydrogen sulphide (H_2S)? $2H_2S + 3O_2 \rightarrow 2SO_2 + 2H_2O$

c) What mass of potassium oxide is formed when 7.8 g of potassium is burned in excess oxygen?
$4K + O_2 \rightarrow 2K_2O$

d) What volume of oxygen (in dm^3) at 298 K and 100 kPa is required to react with 10 g of ammonia?
$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$

e) What mass of aluminium oxide is produced when 135 g of aluminium is burned in oxygen?
$2Al + 3O_2 \rightarrow Al_2O_3$

f) What mass of iodine is produced when 2.4 dm^3 of chlorine gas reacts with excess potassium iodide at 298 K and 100 kPa? $Cl_2 + 2KI \rightarrow 2KCl + I_2$

g) What volume (in dm^3) of hydrogen is needed to react with 32 g of copper oxide at 200 ^oC and 100 kPa?
$CuO + H_2 \rightarrow Cu + H_2O$

h) What volume of oxygen is formed at 398 K and 100 kPa when 735 g of potassium chlorate decomposes?
$2KClO_3 \rightarrow 2KCl + 3O_2$

i) What volume of hydrogen is produced when 195 g of potassium is added to water at 298 K and 100 kPa?
$2K + 2H_2O \rightarrow 2KOH + H_2$

j) What mass of calcium carbonate is required to produce 1.2 dm^3 of carbon dioxide at 398 K and 100 kPa?
$CaCO_3 \rightarrow CaO + CO_2$

k) What mass of magnesium oxide is formed when magnesium reacts with 6 dm^3 of oxygen at 298 K and 100 kPa? $2Mg + O_2 \rightarrow 2MgO$

l) What volume of carbon dioxide (in dm^3) is produced when 5.6 g of butene (C_4H_8) is burnt at 298 K and 100 kPa? $C_4H_8 + 6O_2 \rightarrow 4CO_2 + 4H_2O$

m) The pollutant sulphur dioxide can be removed from the air by reaction with calcium carbonate in the presence of oxygen. What mass of calcium carbonate is needed to remove 480 dm^3 of sulphur dioxide at 298 K and 100 kPa? $2CaCO_3 + 2SO_2 + O_2 \rightarrow 2CaSO_4 + 2CO_2$

n) 25 cm^3 of a solution of sodium hydroxide reacts with 15 cm^3 of 0.1 $moldm^{-3}$ HCl. What is the molar concentration of the sodium hydroxide solution? $HCl + NaOH \rightarrow NaCl + H_2O$

o) Calculate the mass of H_2O required to react completely with 5.0 g of $SiCl_4$:
$SiCl_4 + 2H_2O \rightarrow SiO_2 + 4HCl$

p) Calculate the mass of phosphorus required to make 200 g of phosphine, PH_3, by the reaction:
$P_4(s) + 3NaOH(aq) + 3H_2O(l) \rightarrow 3NaH_2PO_2(aq) + PH_3(g)$

q) Lead (IV) oxide reacts with concentrated hydrochloric acid as follows:
$PbO_2(s) + 4HCl(aq) \rightarrow PbCl_2(s) + Cl_2(g) + 2H_2O(l)$
What mass of lead chloride would be obtained from 37.2g of PbO_2?

r) When copper (II) nitrate is heated, it decomposes according to the following equation:
$2Cu(NO_3)_2(s) \rightarrow 2CuO(s) + 4NO_2(g) + O_2(g)$.
If 20.0g of copper (II) nitrate is heated, what mass of copper (II) oxide would be produced?

s) 25 cm^3 of a solution of 0.1 $moldm^{-3}$ NaOH reacts with 50 cm^3 of a solution of hydrochloric acid (HCl). What is the molarity of the acid? $HCl + NaOH \rightarrow NaCl + H_2O$

t) 25.0 cm^3 of a 0.10 $moldm^{-3}$ solution of sodium hydroxide was titrated against a solution of hydrochloric acid of unknown concentration. 27.3 cm^3 of the acid was required. What was the concentration of the acid?
$HCl + NaOH \rightarrow NaCl + H_2O$

Extension 10.2: Calculating Reacting Quantities

a) 10 cm^3 of a solution of NaCl react with 15 cm^3 of a 0.02 $moldm^{-3}$ solution of $AgNO_3$. What is the concentration of the NaCl solution in gdm^{-3}? $NaCl + AgNO_3 + AgCl + Na\ NO_3$

b) 25 cm^3 of a 0.1 $moldm^{-3}$ solution of an acid H_xA reacts with 75 cm^3 of a 0.1 $moldm^{-3}$ solution of NaOH. What is the value of x? $H_xA + xNaOH \rightarrow + Na_xA + xH_2O$

c) A solution of hydrochloric acid (HCl) of volume 25.0 cm^3 was pipetted onto a piece of marble (calcium carbonate). When all action had ceased, 1.30g of the marble had dissolved. Find the concentration of the acid. $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$

d) What volume of 0.1 $moldm^{-3}$ hydrochloric acid (HCl) would be required to dissolve 2.3 g of calcium carbonate? $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

e) 2.05 g of the carbonate of an unknown alkali metal (X_2CO_3) required 8.9 cm^3 of 2.0 $moldm^{-3}$ hydrochloric acid to completely dissolve it. What was the relative atomic mass of the metal and which metal was it?

$$X_2CO_3(s) + 2HCl(aq) \rightarrow 2XCl(aq) + CO_2(g) + H_2O(l)$$

f) 10.0 g of calcium nitrate is heated at 100 kPa and a temperature of 300 °C, at which temperature it fully decomposes according to the equation: $2Ca(NO_3)_2(s) \rightarrow 2CaO(s) + 4NO_2(g) + O_2(g)$; Calculate

(i) the volume of nitrogen dioxide evolved
(ii) the volume of oxygen evolved
(iii) the total volume of gas evolved

g) Calculate the volume of oxygen produced at 298 K and 100 kPa by the decomposition of 30 cm^3 of 0.1 $moldm^{-3}$ hydrogen peroxide. $2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$

h) Lead (IV) oxide dissolves in concentrated hydrochloric acid according to the following equation: $PbO_2(s) + 4HCl(aq) \rightarrow PbCl_2(s) + Cl_2(g) + 2H_2O(l)$; starting with 37.2 g of lead (IV) oxide, calculate:

(i) the volume of 12 $moldm^{-3}$ HCl needed to completely dissolve it
(ii) the mass of $PbCl_2$ produced
(iii) the volume of chlorine produced at 298 K and 100 kPa

i) What mass of magnesium, and what volume of 2.0 $moldm^{-3}$ hydrochloric acid, will be required to produce 100 cm^3 of hydrogen gas at 298 K and 100 kPa? $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

j) 0.52 g of sodium was added to 100 cm^3 of water. The following reaction takes place: $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$; calculate:

(i) The volume of hydrogen evolved at 298 K and 100 kPa
(ii) The concentration of the sodium hydroxide solution produced, assuming the volume of water does not change.

Lesson 11 – What have I understood about Amount of Substance and Measurement?

11.1 END-OF-TOPIC QUIZ
UNIT 3 – AMOUNT OF SUBSTANCE AND MEASUREMENT

1. List three base quantities commonly measured in Chemistry and give the SI units of each

2. List two derived quantities commonly measured in Chemistry and give the SI units and the base units of each

3. What instruments are used in Chemistry to measure temperature, time and gas volume?

4. A hydrocarbon containing 82.8% carbon and 17.2% hydrogen by mass has a relative molecular mass of 58; deduce its empirical formula and its molecular formula

5. An ionic compound is found to contain 41.7% magnesium, 54.9% oxygen and 3.4% hydrogen by mass; deduce its empirical formula and its unit formula

6. Calculate the mass of 65% HNO_3 required to make 250 cm^3 of a 1.0 $moldm^{-3}$ solution, and explain briefly how you would carry out the procedure

7. Calculate the volume of 6 $moldm^{-3}$ NaOH required to make 250 cm^3 of a 0.5 $moldm^{-3}$ solution

8. N_2 reacts with H_2 to make NH_3. Write a balanced equation for this reaction

9. Use your equation in Q8 to deduce the number of moles of NH_3 which can be made from 0.06 moles of H_2

10. Calculate the volume of H_2 produced at 100 kPa and 298 K when 2.0 g of Na is added to water
($2Na + 2H_2O \rightarrow 2NaOH + H_2$)

11. Calculate the volume of 1.5 $moldm^{-3}$ HNO_3 required to dissolve 4.0 g of $CaCO_3$
($CaCO_3 + 2NHO_3 \rightarrow Ca(NO_3)_2 + CO_2 + H_2O$)

12. Calculate the volume of Cl_2 produced at 100 kPa and 298 K when excess MnO_2 is added to 20 cm^3 of 12 $moldm^{-3}$ HCl ($MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + 2H_2O$)