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Topic 1 Stoichiometric relationships

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Physical and chemical changes

Physical changes

In a physical change, no new substances are produced.

The melting of ice is a physical change.
No new substances are produced.

$$H_2O_{(s)} \rightarrow H_2O_{(l)}$$

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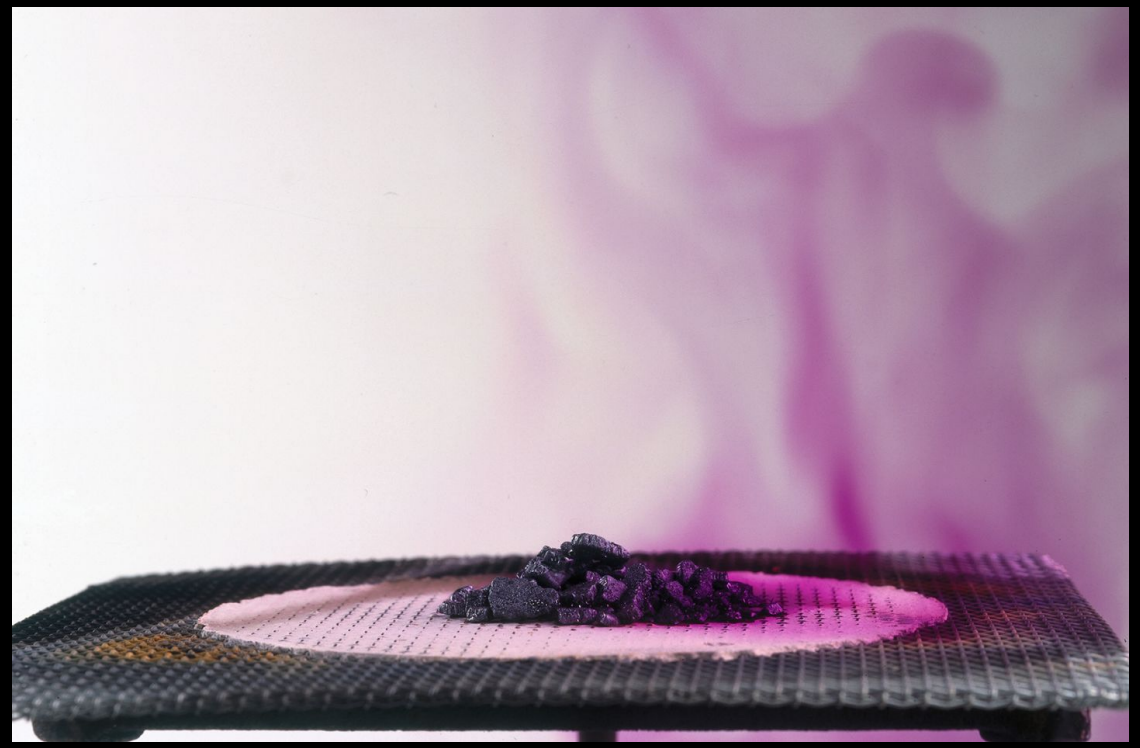
Physical changes

Evaporation

$$Br_{2(l)} \rightarrow Br_{2(g)}$$

Sublimation



$$I_{2(s)} \rightarrow I_{2(g)}$$

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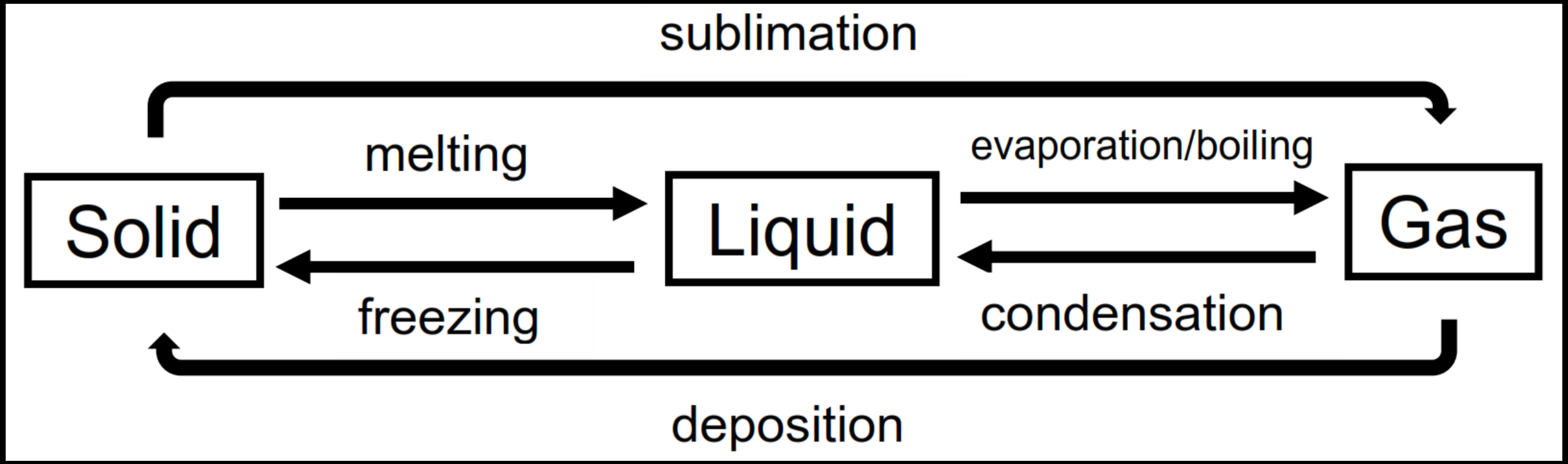
Physical changes

Heat is absorbed (endothermic)



Heat is released (exothermic)

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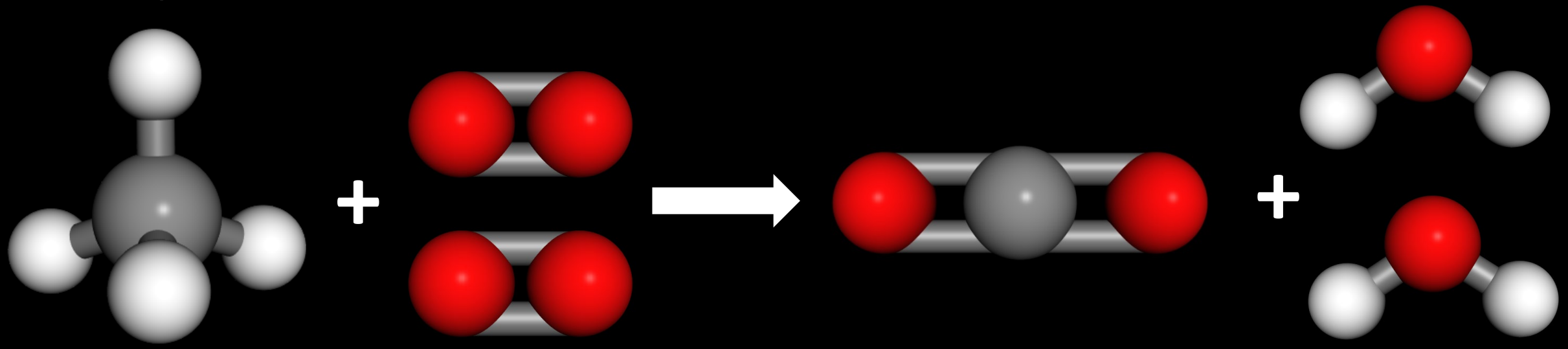
Chemical changes

A chemical change results in the formation of new chemical substances.

In a chemical reaction, atoms are rearranged to form new products.



$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$$

Chemical changes

In a chemical change, the products often have very different properties from the reactants.

$$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$$

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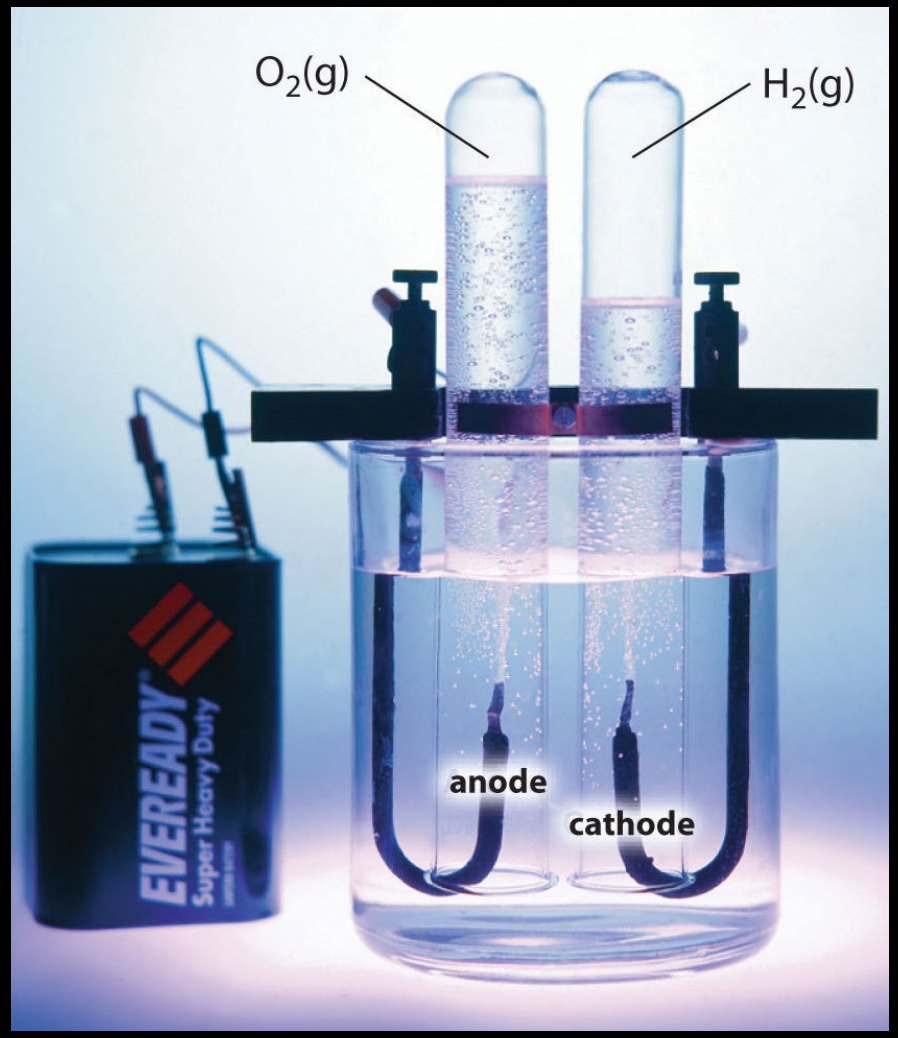
Chemical changes

Electrolysis



Rusting

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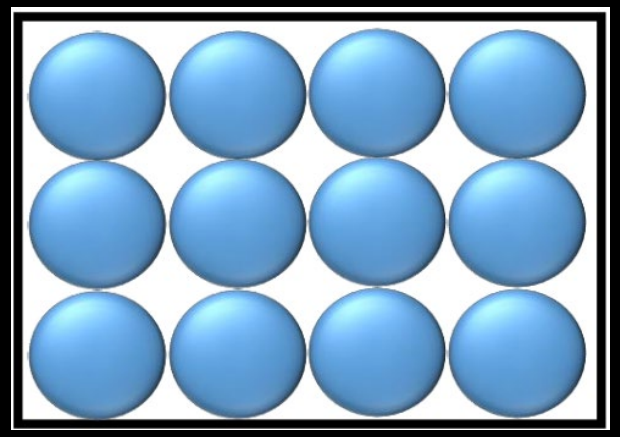
States of matter

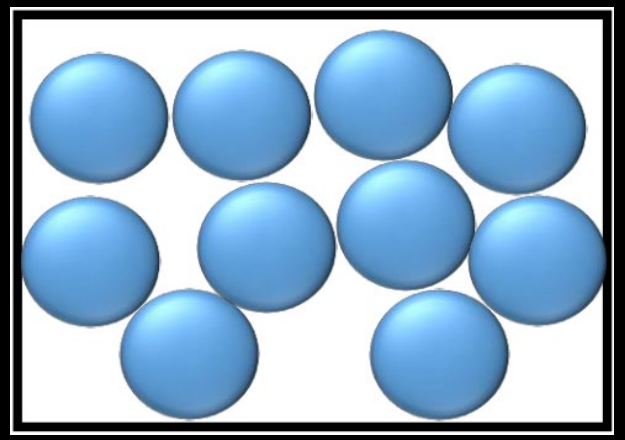
States of matter

Solid



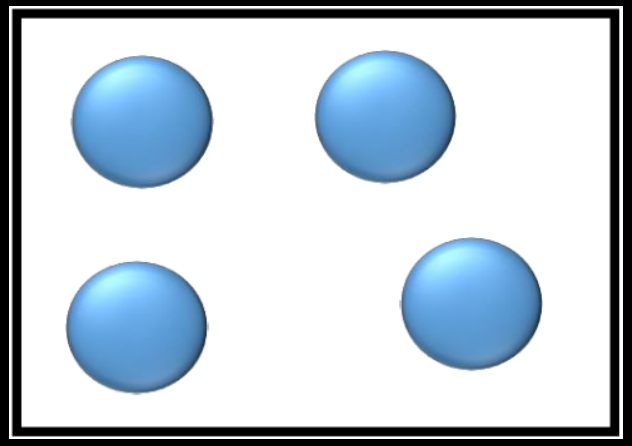
Solids have a fixed shape and volume

Liquid



Liquids have a fixed volume but no fixed shape

Gas



Gases have neither a fixed volume nor a fixed shape

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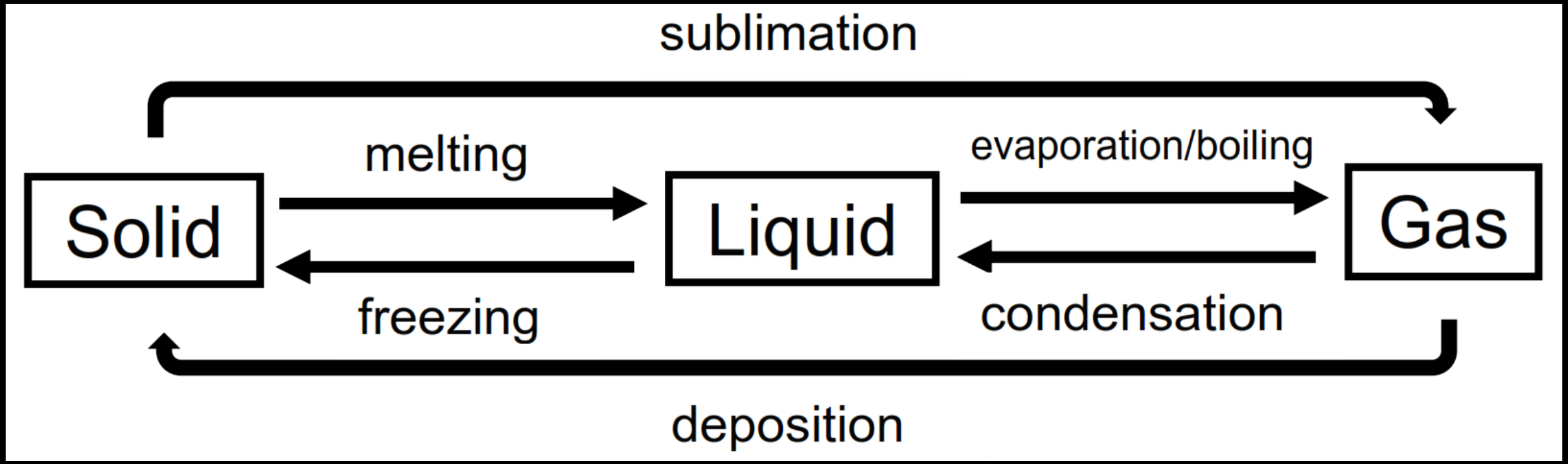
Physical changes

Heat is absorbed (endothermic)



Heat is released (exothermic)

State symbols

State symbols show the physical state of a substance.

(s) – solid

(l) – liquid

(g) – gas

(aq) – aqueous

$$2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$$

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Elements, compounds and mixtures

Matter

Matter

- Pure substances
  - Element and compounds
- Mixtures
  - Homogeneous or heterogeneous

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Elements

An element is a substance that cannot be broken down into a simpler substance by chemical means.

6. The periodic table

Atomic number
Element
Relative atomic mass

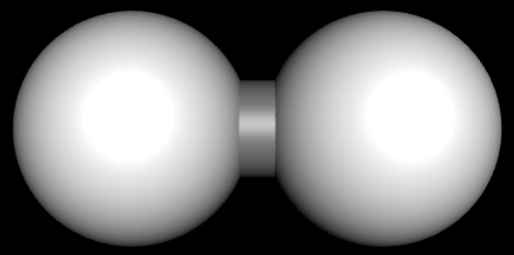
	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1.01																	2 He 4.00
2	3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.90
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.96	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
6	55 Cs 132.91	56 Ba 137.33	57 † La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.20	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	89 ‡ Ac (227)	104 Rf (267)	105 Db (268)	106 Sg (269)	107 Bh (270)	108 Hs (269)	109 Mt (278)	110 Ds (281)	111 Rg (281)	112 Cn (285)	113 Uut (286)	114 Uuq (289)	115 Uup (288)	116 Uuh (293)	117 Uus (294)	118 Uuo (294)

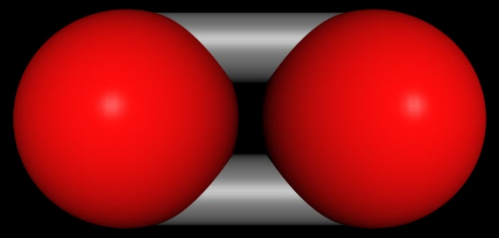
†	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.05	71 Lu 174.97
‡	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

Elements

Some elements exist as diatomic molecules.
A molecule is an electrically neutral group of two or more atoms bonded together.

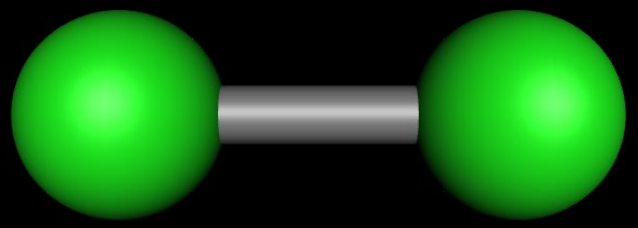
Hydrogen
H_2

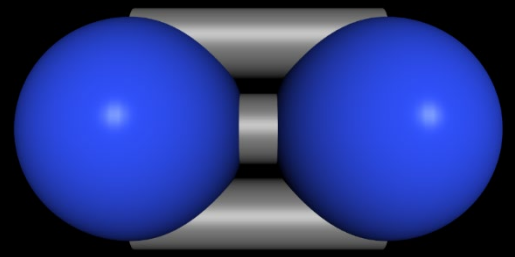




Oxygen O_2

Chlorine
Cl_2

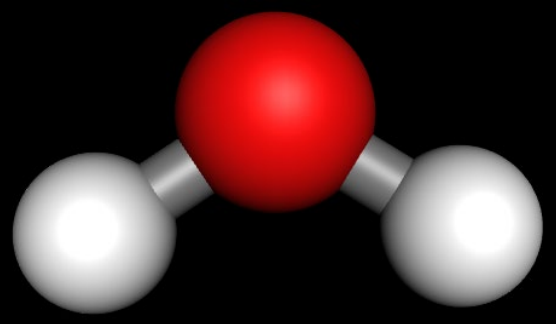




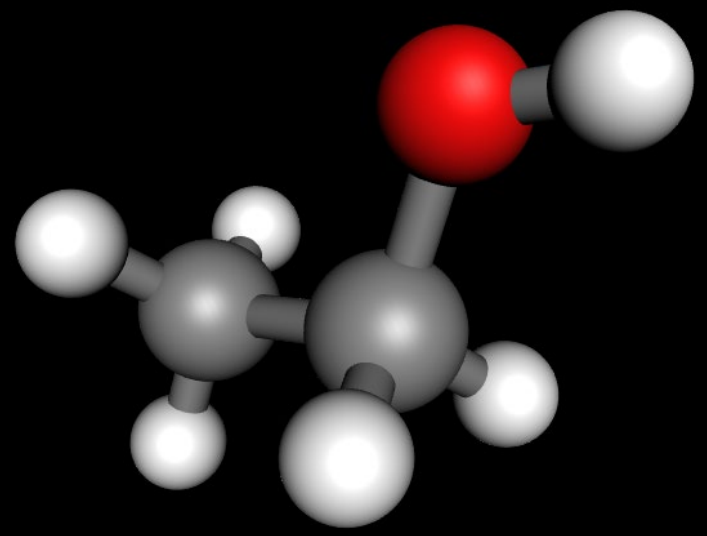
Nitrogen N_2

Compounds

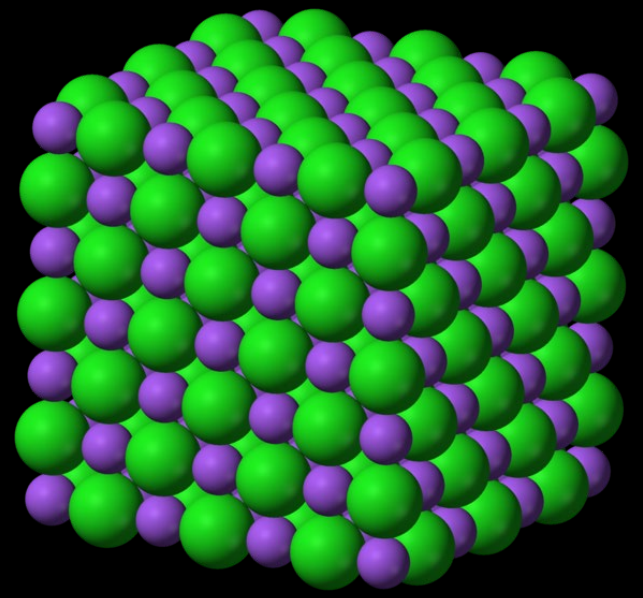
A compound is formed from two or more different elements chemically joined in a fixed ratio.



Water H_2O



Ethanol
C_2H_5OH



Sodium chloride
NaCl

Compounds

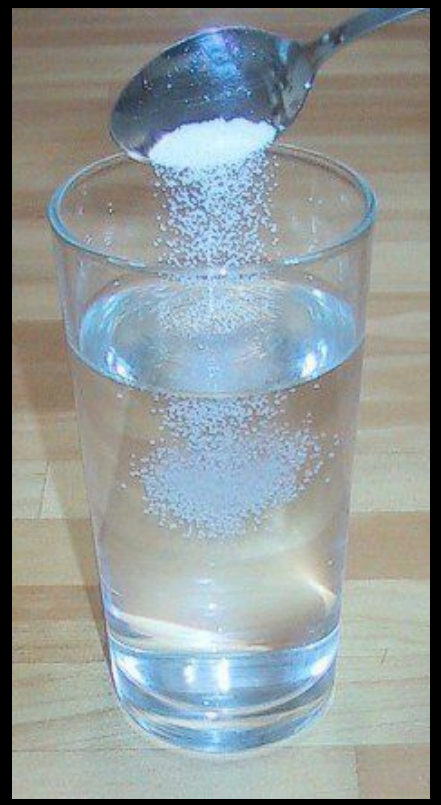
Compounds have different properties from the elements that they are made from.

$$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$$

Mixtures

Mixtures contain more than one element and/or compound that are not chemically bonded together and so retain their individual properties.

Homogeneous mixtures have a constant composition throughout.



Heterogeneous mixtures have visibly different substances or phases.

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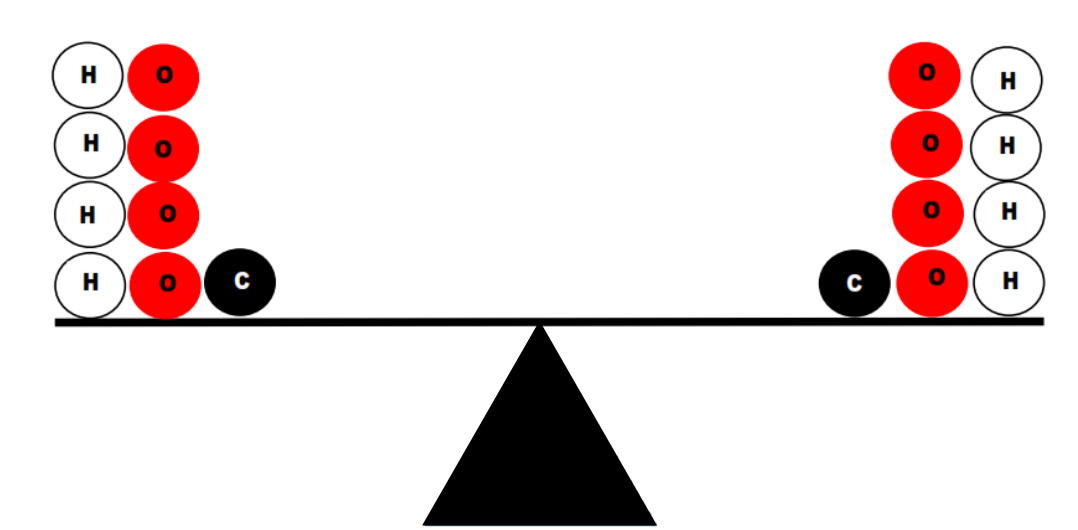
Balancing chemical equations

Balancing chemical equations

The law of the conservation of mass states that mass is conserved in a chemical reaction.

The total mass of the products must be equal to the total mass of the reactants.

1 × C (12.01 g)
4 × H (4.04 g)
4 × O (64.00 g)
Total = 80.05 g



$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}$

1 × C (12.01 g)
4 × H (4.04 g)
4 × O (64.00 g)
Total = 80.05 g

Balancing chemical equations

To balance a chemical equation, we can only change the numbers in front of the reactants and products – these are called coefficients.

$$Na_{(s)} + Cl_{2(g)} \rightarrow NaCl_{(s)}$$

Na	1	2	Na	1	2
Cl	2		Cl	1	2

$$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$$

Balancing chemical equations

$$CaCO_{3(s)} + HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

Ca 1		Ca 1
C 1		C 1
H 1	2	H 2
O 3		O 3
Cl 1	2	Cl 2

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

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Balancing chemical equations

$$C_6H_{14(l)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$$

C	6	12	C	1	12
H	14	28	H	2	28
O	2	38	O	3	38

$$2C_6H_{14(l)} + 19O_{2(g)} \rightarrow 12CO_{2(g)} + 14H_2O_{(l)}$$

Balancing chemical equations

$$Zn_{(s)} + AgNO_{3(aq)} \rightarrow Zn(NO_3)_{2(aq)} + Ag_{(s)}$$

Zn	1		Zn	1
Ag	1	2	Ag	1 2
N	1	2	N	2
O	3	6	O	6

$$Zn_{(s)} + 2AgNO_{3(aq)} \rightarrow Zn(NO_3)_{2(aq)} + 2Ag_{(s)}$$

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Relative atomic mass and molecular mass

Relative atomic mass (A_r)

Relative atomic mass, A_r, is the weighted average mass of the naturally occurring isotopes of an element relative to 1/12 the mass of an atom of carbon-12.

$^{12}_{6}C$

The relative atomic mass scale is based on the isotope carbon-12 which has a mass of exactly 12 amu.

Atomic number	1	12	17	26
Element	**H**	**Mg**	**Cl**	**Fe**
Relative atomic mass	1.01	24.31	35.45	55.85

Relative atomic mass (A_r)

Element	Relative atomic mass	Mass compared to ^{12}C
Hydrogen	1.01	≈ 12 times lighter
Helium	4.00	≈ 3 times lighter
Magnesium	24.31	≈ 2 times heavier
Phosphorus	30.07	≈ 2.5 times heavier
Chlorine	35.45	≈ 3 times heavier

Relative atomic mass (A_r)

Isotope	Percent abundance (%)
^{24}Mg	78.99
^{25}Mg	10.00
^{26}Mg	11.01

$$A_r = \frac{(24 \times 78.99) + (25 \times 10.00) + (26 \times 11.01)}{100}$$

$$A_r = 24.32$$

Relative molecular mass (M_r)

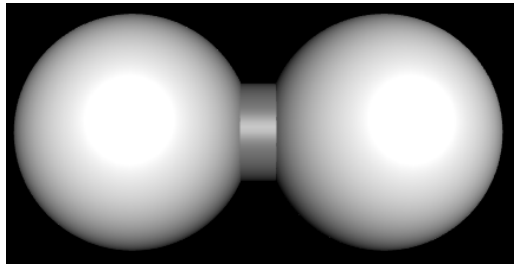
Relative molecular mass, M_r, is the weighted average mass of a molecule relative to 1/12 the mass of an atom of ^{12}C. The M_r is the sum of the A_r of the atoms in the molecule.

Molecule	Atoms	Relative molecular mass
H_2	2 × H (1.01)	2.02
H_2O	2 × H (1.01) 1 × O (16.00)	18.02
C_2H_6	2 × C (12.01) 6 × H (1.01)	30.08

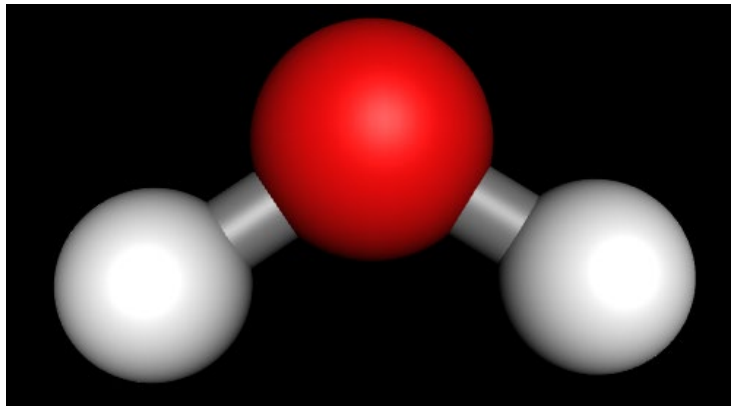
Relative molecular mass (M_r)

Relative molecular mass, M_r, is the weighted average mass of a molecule relative to 1/12 the mass of an atom of carbon-12.

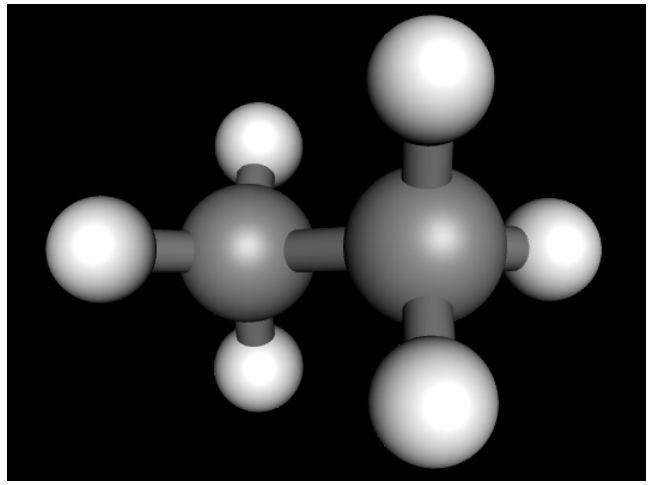
It is the sum of the A_r of the atoms in the substance.



H_2 M_r = 2.02



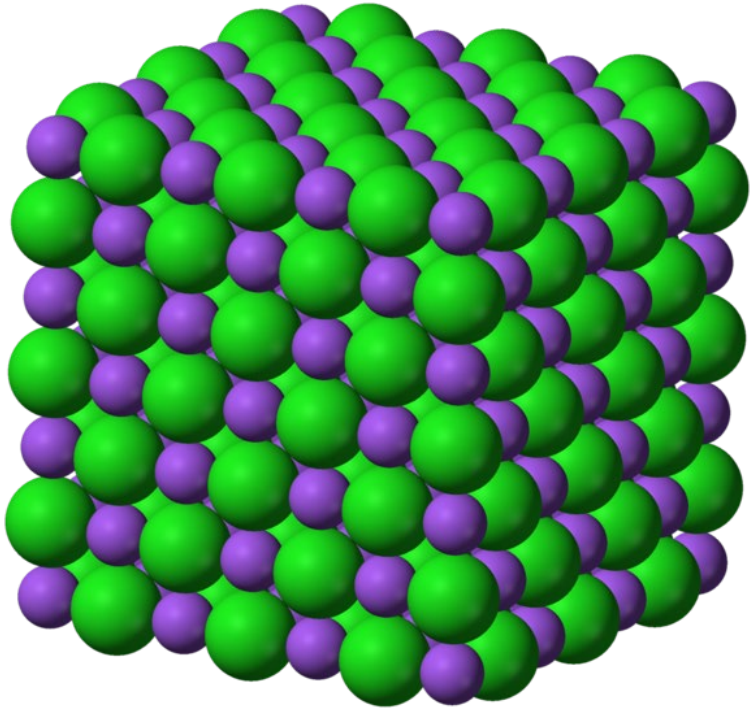
H_2O M_r = 18.02



C_2H_6 M_r = 30.08

Relative formula mass (M_r)

Relative formula mass is mostly used for compounds that do not form molecules, such as ionic compounds.



11	17
Na	**Cl**
22.99	35.45

The relative formula mass of sodium chloride, NaCl, is 58.44

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Molar mass

Molar mass

Molar mass (*M*) is the mass in grams of one mole of a substance (g mol^{-1}).

One mole of substance contains 6.02×10^{23} particles.

The molar mass of a substance is numerically equal to its relative atomic mass.

Atomic number	6	12	16	26
Element	**C**	**Mg**	**S**	**Fe**
Relative atomic mass	12.01	24.31	32.07	55.85

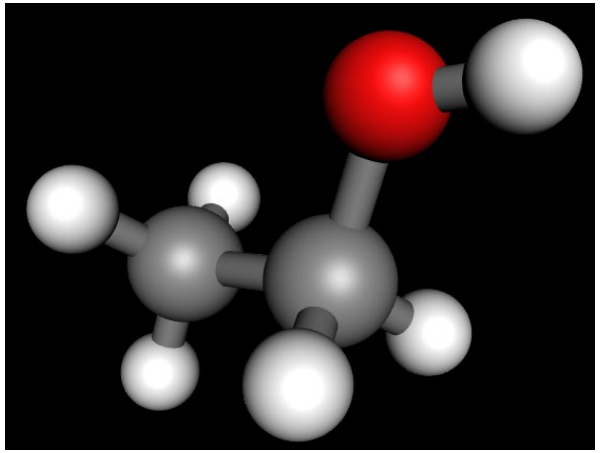
Molar mass

To convert A_r to M, multiply by the molar mass constant, M_u, which is approximately equal to 1 g mol^{-1}

Element	Relative atomic mass	Molar mass (g mol^{-1})
C	12.01	12.01
Mg	24.31	24.31
S	32.07	32.07
Fe	55.85	55.85

Molar mass

Determine the molar mass of ethanol, C_2H_5OH.



2 carbon atoms $A_r = 12.01$

1 oxygen atom $A_r = 16.00$

6 hydrogen atoms $A_r = 1.01$

$M_r = (2 \times 12.01) + 16.00 + (6 \times 1.01) = 46.08$

$M = 46.08 \times M_u$ (≈ 1 g mol^{-1})

$M = 46.08$ g mol^{-1}

Molar mass

Substance	Relative molecular mass/formula mass	Molar mass *M* (g mol^{-1})
O_2	32.00	32.00
H_2O	18.02	18.02
CH_4	16.05	16.05
NaCl	58.44	58.44
$(NH_4)_2CO_3$	96.11	96.11
Al_2O_3	101.96	101.96

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The mole concept

The mole concept

How many carbon atoms are there in 1.00 mol of ethanol, C_2H_5OH?

How many hydrogen atoms are there in 2.00 mol of methane, CH_4?

How many Na^+ ions are there in 1.00 mol of NaCl?

What is the total number of ions in 0.50 mol of $(NH_4)_2CO_3$?

The mole concept

The mole, symbol mol, is the SI unit for amount of substance (*n*).

One mole contains exactly $6.02214076 \times 10^{23}$ elementary entities.

$$6.02 \times 10^{23}$$

The Avogadro constant, *L* or N_A is:

$$6.02 \times 10^{23}\ \text{mol}^{-1}$$

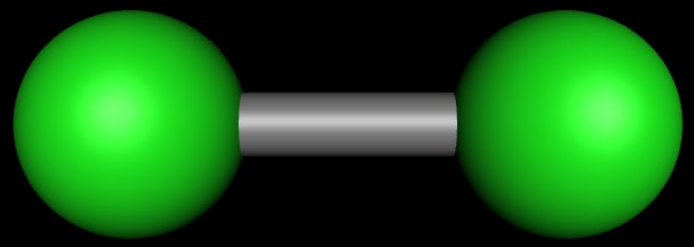
The mole concept

Elementary entity	Number of elementary entities in one mole
Atoms	6.02×10^{23}
Molecules	6.02×10^{23}
Ions	6.02×10^{23}
Formula units	6.02×10^{23}

The mole concept

Determine the number of chlorine molecules and chlorine atoms in 1.00 mol of chlorine gas, Cl_2.

$6.02 \times 10^{23} \times$

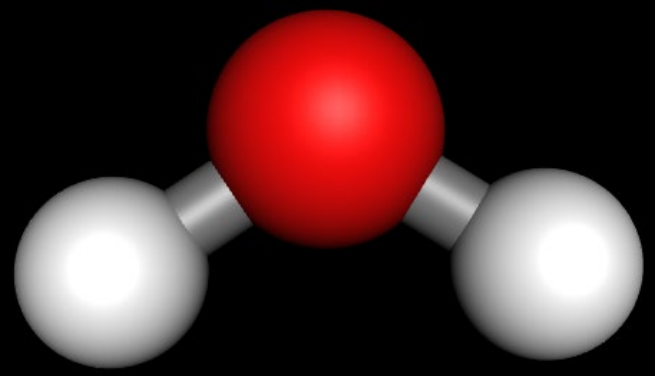


Cl_2 molecules	6.02×10^{23}
Cl atoms	$2 \times 6.02 \times 10^{23} = 1.20 \times 10^{24}$

The mole concept

Determine the number of hydrogen atoms and oxygen atoms in 0.500 mol of water, H_2O.

$$0.500 \times 6.02 \times 10^{23} \times$$

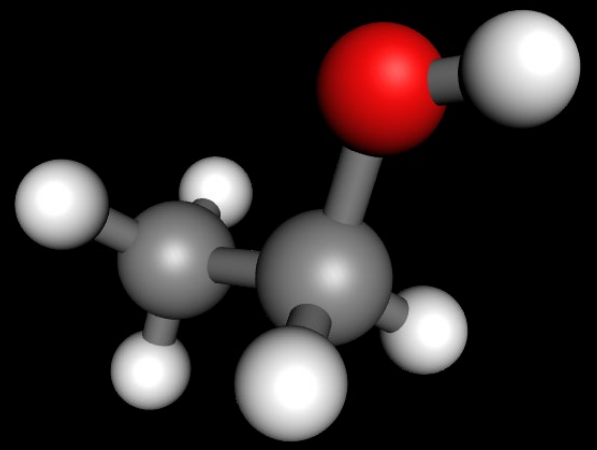


H atoms	$2 \times 0.500 \times 6.02 \times 10^{23} = 6.02 \times 10^{23}$
O atoms	$0.500 \times 6.02 \times 10^{23} = 3.01 \times 10^{23}$

The mole concept

Determine the number of carbon atoms, hydrogen atoms and oxygen atoms in 0.250 mol of ethanol, C_2H_5OH.

$$0.250 \times 6.02 \times 10^{23} \times$$

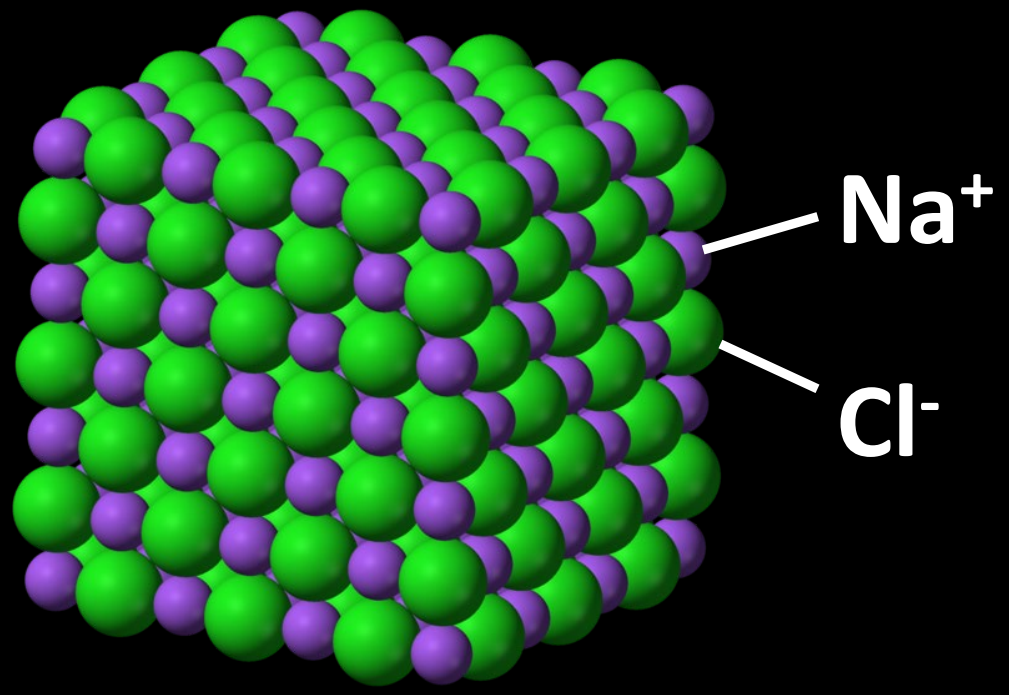


C atoms	$2 \times 0.250 \times 6.02 \times 10^{23} = 3.01 \times 10^{23}$
H atoms	$6 \times 0.250 \times 6.02 \times 10^{23} = 9.03 \times 10^{23}$
O atoms	$0.250 \times 6.02 \times 10^{23} = 1.51 \times 10^{23}$

The mole concept

In one mole of sodium chloride (NaCl) there are 6.02×10^{23} NaCl formula units.

NaCl is an ionic compound therefore it does not form molecules.



One mole of NaCl has 6.02×10^{23} sodium ions and 6.02×10^{23} chloride ions (total number of ions = 1.20×10^{24}).

The mole concept

Determine the number of magnesium ions and chloride ions in 1.00 mol of magnesium chloride, $MgCl_2$.

$$6.02 \times 10^{23} \times MgCl_2$$

Mg^{2+} ions	6.02×10^{23}
Cl^- ions	$2 \times 6.02 \times 10^{23} = 1.20 \times 10^{24}$

The mole concept

Determine the number of protons, neutrons and electrons in 0.750 mol of carbon-12 atoms.

$$0.750 \times 6.02 \times 10^{23} \times {}^{12}_{6}C$$

Protons	$0.750 \times 6 \times 6.02 \times 10^{23} = 2.71 \times 10^{24}$
Electrons	$0.750 \times 6 \times 6.02 \times 10^{23} = 2.71 \times 10^{24}$
Neutrons	$0.750 \times 6 \times 6.02 \times 10^{23} = 2.71 \times 10^{24}$

The mole concept

How many carbon atoms are there in 1.00 mol of ethanol, C_2H_5OH? 1.20×10^{24} C atoms

How many hydrogen atoms are there in 2.00 mol of methane, CH_4? 4.82×10^{24} H atoms

How many Na^+ ions are there in 1.00 mol of NaCl?

6.02×10^{23} Na^+ ions

What is the total number of ions in 0.50 mol of $(NH_4)_2CO_3$? 9.03×10^{23} ions

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Calculating the amount of substance

Calculating amount of substance

How to calculate the amount (in mol) of a substance from its mass (*m*) and molar mass (*M*).

$$\text{amount of substance (mol)} = \frac{\text{mass (g)}}{\text{molar mass (g mol}^{-1}\text{)}}$$

$$n(\text{mol}) = \frac{m\ (\text{g})}{M\ (\text{g mol}^{-1})} \qquad n = \frac{m}{M}$$

Calculating amount of substance

$$\text{mass (g)} = \text{amount (mol)} \times \text{molar mass } (\text{g mol}^{-1})$$

$$m = nM$$

$$\text{molar mass } (\text{g mol}^{-1}) = \frac{\text{mass (g)}}{\text{amount (mol)}}$$

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

Calculating amount of substance

Calculate the amount (in mol) of O_2 in a 16.00 g sample of O_2.

$$M(\mathrm{O_2}) = 16.00 \times 2 = 32.00\ \mathrm{g\ mol^{-1}}$$

$$n(\mathrm{O_2}) = \frac{16.00\ \mathrm{g}}{32.00\ \mathrm{g\ mol^{-1}}}$$

$$n(\mathrm{O_2}) = 0.5000\ \mathrm{mol}$$

Calculating amount of substance

Calculate the amount (in mol) of H_2O in a 100.0 g sample of H_2O.

$$M(\mathrm{H_2O}) = 16.00 + (2 \times 1.01) = 18.02\ \mathrm{g\ mol^{-1}}$$

$$n(\mathrm{H_2O}) = \frac{100.0\ \mathrm{g}}{18.02\ \mathrm{g\ mol^{-1}}}$$

$$n(\mathrm{H_2O}) = 5.549\ \mathrm{mol}$$

Calculating amount of substance

Calculate the amount (in mol) of NaCl in a 50.00 g sample of NaCl.

$$M(\mathrm{NaCl}) = 22.99 + 35.45 = 58.44\ \mathrm{g\ mol^{-1}}$$

$$n(\mathrm{NaCl}) = \frac{50.00\ \mathrm{g}}{58.44\ \mathrm{g\ mol^{-1}}}$$

$$n(\mathrm{NaCl}) = 0.8556\ \mathrm{mol}$$

Calculating amount of substance

Calculate the amount (in mol) of $Ni(NO_3)_2$ in a 75.23 g sample of $Ni(NO_3)_2$.

$$M(\mathrm{Ni(NO_3)_2}) = 58.69 + (2 \times 14.01) + (6 \times 16.00) = 182.71\ \mathrm{g\ mol^{-1}}$$

$$n(\mathrm{Ni(NO_3)_2}) = \frac{75.23\ \mathrm{g}}{182.71\ \mathrm{g\ mol^{-1}}}$$

$$n(\mathrm{Ni(NO_3)_2}) = 0.4117\ \mathrm{mol}$$

Calculating amount of substance

How to calculate the amount (in mol) of a substance from its volume (*V*) and concentration (*c*).

$$\text{amount (mol)} = \text{concentration}\left(\text{mol dm}^{-3}\right) \times \text{volume}\left(\text{dm}^3\right)$$

$$n(\text{mol}) = c\left(\text{mol dm}^{-3}\right) \times V\left(\text{dm}^3\right)$$

$$n = cV$$

$$1\ \text{cm}^3 \times \frac{1\ \text{dm}^3}{1000\ \text{cm}^3} = 0.001\ \text{dm}^3$$

Calculating amount of substance

$$\text{concentration}\left(\text{mol dm}^{-3}\right) = \frac{\text{amount (mol)}}{\text{volume (dm}^3\text{)}}$$

$$c = \frac{n}{V}$$

$$\text{volume}\left(\text{dm}^3\right) = \frac{\text{amount (mol)}}{\text{concentration (mol dm}^{-3}\text{)}}$$

$$V = \frac{n}{c}$$

Calculating amount of substance

Calculate the amount (in mol) of HCl in 100.0 cm^3 of 0.500 mol dm^{-3} $HCl_{(aq)}$

$$100.0\ \mathrm{cm}^3 \times \frac{1\ \mathrm{dm}^3}{1000\ \mathrm{cm}^3} = 0.100\ \mathrm{dm}^3$$

$$n(\mathrm{HCl}) = 0.500\ \mathrm{mol\ dm}^{-3} \times 0.100\ \mathrm{dm}^3$$

$$n(\mathrm{HCl}) = 0.0500\ \mathrm{mol}$$

Calculating amount of substance

Calculate the amount (in mol) of NaOH in 50.0 cm^3 of 2.00 mol dm^{-3} $NaOH_{(aq)}$

$$50.0\ \mathrm{cm}^3 \times \frac{1\ \mathrm{dm}^3}{1000\ \mathrm{cm}^3} = 0.0500\ \mathrm{dm}^3$$

$$n(\mathrm{NaOH}) = 2.00\ \mathrm{mol\ dm}^{-3} \times 0.0500\ \mathrm{dm}^3$$

$$n(\mathrm{NaOH}) = 0.100\ \mathrm{mol}$$

Calculating amount of substance

Calculate the amount (in mol) of NaCl in 60.0 cm^3 of 0.850 mol dm^{-3} $NaCl_{(aq)}$.

$$60.0\ \mathrm{cm}^3 \times \frac{1\ \mathrm{dm}^3}{1000\ \mathrm{cm}^3} = 0.0600\ \mathrm{dm}^3$$

$$n(\mathrm{NaCl}) = 0.850\ \mathrm{mol\ dm}^{-3} \times 0.0600\ \mathrm{dm}^3$$

$$n(\mathrm{NaCl}) = 0.0510\ \mathrm{mol}$$

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Calculating mass (g) from amount (in mol)

Calculating mass from amount

$$n(\text{mol}) = \frac{m\ (\text{g})}{M\ (\text{g mol}^{-1})} \qquad n = \frac{m}{M}$$

$$\text{mass (g)} = \text{amount (mol)} \times \text{molar mass } (\text{g mol}^{-1})$$

$$m = nM$$

Calculating mass from amount

Calculate the mass (in g) of 0.6437 mol of $CaCO_3$.

$$M(\mathrm{CaCO_3}) = 40.08 + 12.01 + (3 \times 16.00)$$

$$M(\mathrm{CaCO_3}) = 100.09\ \mathrm{g\ mol^{-1}}$$

$$m = nM$$

$$m = 0.6437\ \mathrm{mol} \times 100.09\ \mathrm{g\ mol^{-1}}$$

$$m = 64.43\ \mathrm{g}$$

Calculating mass from amount

Calculate the mass (in g) of 0.8539 mol of $AlCl_3$

$$M(\mathrm{AlCl_3}) = 26.98 + (3 \times 35.45)$$

$$M(\mathrm{AlCl_3}) = 133.33\ \mathrm{g\ mol^{-1}}$$

$$m = nM$$

$$m = 0.8539\ \mathrm{mol} \times 133.33\ \mathrm{g\ mol^{-1}}$$

$$m = 113.9\ \mathrm{g}$$

Calculating mass from amount

Calculate the mass (in g) of 1.379 mol of $C_6H_{12}O_6$

$$M(\mathrm{C_6H_{12}O_6}) = (6 \times 12.01) + (12 \times 1.01) + (6 \times 16.00)$$

$$M(\mathrm{C_6H_{12}O_6}) = 180.18\ \mathrm{g\ mol^{-1}}$$

$$m = nM$$

$$m = 1.379\ \mathrm{mol} \times 180.18\ \mathrm{g\ mol^{-1}}$$

$$m = 248.5\ \mathrm{g}$$

Calculating mass from amount

Calculate the mass (in g) of 1.264 mol of $Ni(NO_3)_2$

$$M(\mathrm{Ni(NO_3)_2}) = 58.69 + (2 \times 14.01) + (6 \times 16.00)$$

$$M(\mathrm{Ni(NO_3)_2}) = 182.71\ \mathrm{g\ mol^{-1}}$$

$$m = nM$$

$$m = 1.264\ \mathrm{mol} \times 182.71\ \mathrm{g\ mol^{-1}}$$

$$m = 230.9\ \mathrm{g}$$

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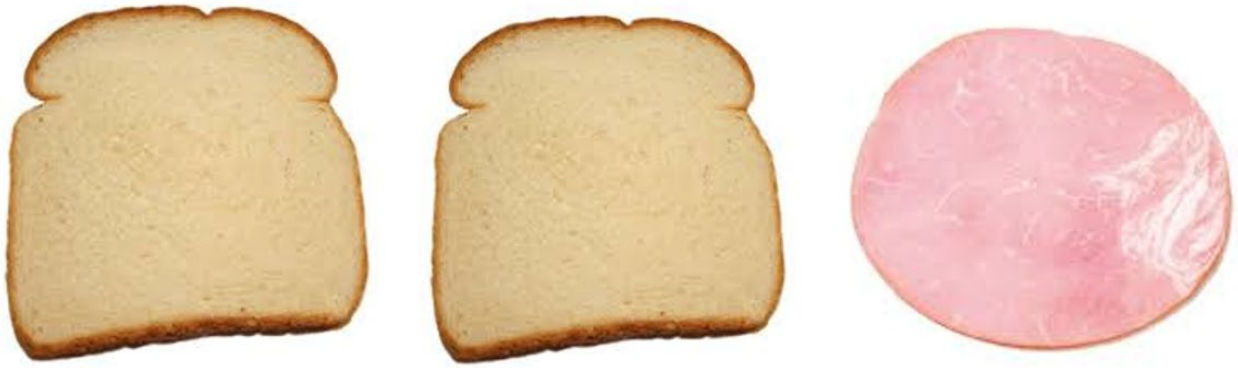
Limiting and excess reactants

Limiting and excess reactants



How many sandwiches can be made with 12 pieces of bread and 7 slices of ham?

6 sandwiches = 12 pieces of bread and 6 slices of ham.

The bread is limiting the number of sandwiches that can be made (the limiting reactant).

One slice of ham remains (the excess reactant).

Limiting and excess reactants

The limiting reactant (reagent) is the reactant that limits the amount of product(s) that can be made. The excess reactant is the reactant that remains when the limiting reactant is consumed.

1. Calculate the amount (in mol) of each reactant.
2. Divide the amount of each reactant by its coefficient in the balanced equation.
3. The lowest value is the limiting reactant.

Limiting and excess reactants

50.0 g of N_2H_4 is reacted with 75.0 g of N_2O_4. Determine the limiting and excess reactants.

$$2N_2H_{4(l)} + N_2O_{4(l)} \rightarrow 3N_{2(g)} + 4H_2O_{(g)}$$

$$n(N_2H_4) = \frac{50.0\ \text{g}}{32.06\ \text{g mol}^{-1}} = 1.56\ \text{mol}$$

$$n(N_2O_4) = \frac{75.0\ \text{g}}{92.02\ \text{g mol}^{-1}} = 0.815\ \text{mol}$$

Limiting and excess reactants

$$2N_2H_{4(l)} + N_2O_{4(l)} \rightarrow 3N_{2(g)} + 4H_2O_{(g)}$$

$$N_2H_4 = \frac{1.56}{2} = 0.780$$

$$N_2O_4 = \frac{0.815}{1} = 0.815$$

N_2H_4 is the limiting reactant and N_2O_4 is the excess reactant.

Limiting and excess reactants

Determine the amount of excess reactant that remains at the end of the reaction.

$$2N_2H_{4(l)} + N_2O_{4(l)} \rightarrow 3N_{2(g)} + 4H_2O_{(g)}$$

Molar ratio of N_2H_4 to N_2O_4 is 2:1

$$1.56 \text{ mol } N_2H_4 \times \frac{1 \text{ mol } N_2O_4}{2 \text{ mol } N_2H_4} = 0.780 \text{ mol } N_2O_4$$

$$n(N_2O_4) \text{ remaining} = 0.815 - 0.780 = 0.0350 \text{ mol}$$

Limiting and excess reactants

Determine the maximum amount of N_2 that can be produced in the reaction.

$$2N_2H_{4(l)} + N_2O_{4(l)} \rightarrow 3N_{2(g)} + 4H_2O_{(g)}$$

Molar ratio of N_2H_4 to N_2 is 2:3

$$1.56 \text{ mol } N_2H_4 \times \frac{3 \text{ mol } N_2}{2 \text{ mol } N_2H_4} = 2.34 \text{ mol } N_2$$

Limiting and excess reactants

3.00 g of Zn is reacted with 50.0 cm^3 of 1.00 mol dm^{-3} HCl. Determine the limiting and excess reactants.

$$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

$$n(\text{Zn}) = \frac{3.00\ \text{g}}{65.38\ \text{g mol}^{-1}} = 0.0459\ \text{mol}$$

$$n(\text{HCl}) = 1.00\ \text{mol dm}^{-3} \times \frac{50.0\ \text{cm}^3}{1000\ \text{cm}^3} = 0.0500\ \text{mol}$$

Limiting and excess reactants

$$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

$$Zn = \frac{0.0459}{1} = 0.0459$$

$$HCl = \frac{0.0500}{2} = 0.0250$$

HCl is the limiting reactant and Zn is the excess reactant.

Limiting and excess reactants

Determine the amount of excess reactant that remains at the end of the reaction.

$$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

Molar ratio of Zn to HCl is 1:2

$$0.0500 \text{ mol HCl} \times \frac{1 \text{ mol Zn}}{2 \text{ mol HCl}} = 0.0250 \text{ mol Zn}$$

$$n(\text{Zn}) \text{ remaining} = 0.0459 - 0.0250 = 0.0209 \text{ mol}$$

Limiting and excess reactants

Determine the maximum amount of H_2 that can be produced in the reaction.

$$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

Molar ratio of HCl to H_2 is 2:1

$$0.0500 \text{ mol HCl} \times \frac{1 \text{ mol } H_2}{2 \text{ mol HCl}} = 0.0250 \text{ mol } H_2$$

Limiting reactant

100.0 g of Fe_2O_3 is reacted with 100.0 g of C. Determine the maximum mass of Fe that can be produced.

$$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$$

$$n(Fe_2O_3) = \frac{100.0}{159.70} = 0.6262 \text{ mol}$$

$$n(C) = \frac{100.0}{12.01} = 8.326 \text{ mol}$$

Limiting reactant

$$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$$

$$Fe_2O_3 \frac{0.6262}{2} = 0.3131$$

$$C \frac{8.326}{3} = 2.775$$

Limiting reactant

$$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$$

$$n(Fe) = 0.6262 \times 2 = 1.252 \text{ mol}$$

$$m = nM$$

$$m = 1.252 \times 55.85$$

$$m = 69.92 \text{ g Fe}$$

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Theoretical yield and percent yield

Theoretical yield and percent yield

The theoretical yield is the maximum amount of product that can be formed in a chemical reaction (based on the stoichiometry of the reaction).

The actual yield is the actual amount of product that is formed in a chemical reaction.

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\ \%$$

Theoretical yield and percent yield

How to calculate the percent yield:

1. Determine which reactant is the limiting reactant (if not already stated in the question).
2. Use the molar ratio of the limiting reactant to the product in question to calculate the amount of product formed in the reaction.
3. Convert from amount (in mol) to mass (in g).
4. Divide the actual yield by the theoretical yield and multiply by 100 %.

Theoretical yield and percent yield

100.0 g of iron(II) oxide is reacted with 100.0 g of carbon. 46.73 g of iron is produced. Calculate the % yield.

$$2Fe_2O_{3(s)} + 3C_{(s)} \rightarrow 4Fe_{(s)} + CO_{2(g)}$$

$$n(Fe_2O_3) = \frac{100.0\ \mathrm{g}}{159.70\ \mathrm{g\ mol^{-1}}} = 0.6262\ \mathrm{mol}$$

$$n(C) = \frac{100.0\ \mathrm{g}}{12.01\ \mathrm{g\ mol^{-1}}} = 8.326\ \mathrm{mol}$$

Theoretical yield and percent yield

$$2Fe_2O_{3(s)} + 3C_{(s)} \rightarrow 4Fe_{(s)} + CO_{2(g)}$$

$$Fe_2O_3 = \frac{0.6262}{2} = 0.3131$$

$$C = \frac{8.326}{3} = 2.775$$

Fe_2O_3 is the limiting reactant and C is the excess reactant.

Theoretical yield and percent yield

$$2Fe_2O_{3(s)} + 3C_{(s)} \rightarrow 4Fe_{(s)} + CO_{2(g)}$$

Molar ratio of Fe_2O_3 to Fe is 2:4

$$0.6262 \text{ mol } Fe_2O_3 \times \frac{4 \text{ mol Fe}}{2 \text{ mol } Fe_2O_3} = 1.252 \text{ mol Fe}$$

$$m(\text{Fe}) = 1.252 \text{ mol} \times 55.85 \text{ g mol}^{-1} = 69.92 \text{ g}$$

Theoretical yield of Fe = 69.92 g

Theoretical yield and percent yield

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\ \%$$

$$\text{Percent yield} = \frac{46.73\ \text{g}}{69.92\ \text{g}} \times 100\ \%$$

$$\text{Percent yield} = 66.83\ \%$$

Theoretical yield and percent yield

15.0 g of $CaCO_3$ is reacted with 50.0 cm^3 of 2.00 mol dm^{-3} HCl. 1.85 g of CO_2 is produced. Calculate the % yield.

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$$

$$n(CaCO_3) = \frac{15.0\ \mathrm{g}}{100.09\ \mathrm{g\ mol^{-1}}} = 0.150\ \mathrm{mol}$$

$$n(HCl) = 2.00\ \mathrm{mol\ dm^{-3}} \times 0.0500\ \mathrm{dm^3} = 0.100\ \mathrm{mol}$$

Theoretical yield and percent yield

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$$

$$CaCO_3 = \frac{0.150}{1} = 0.150$$

$$HCl = \frac{0.100}{2} = 0.0500$$

HCl is the limiting reactant and $CaCO_3$ is the excess reactant.

Theoretical yield and percent yield

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$$

Molar ratio of HCl to CO_2 is 2:1

$$0.100 \text{ mol HCl} \times \frac{1 \text{ mol } CO_2}{2 \text{ mol HCl}} = 0.0500 \text{ mol } CO_2$$

$$m(CO_2) = 0.0500 \text{ mol} \times 44.01 \text{ g mol}^{-1} = 2.20 \text{ g}$$

Theoretical yield of CO_2 = 2.20 g

Theoretical yield and percent yield

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\ \%$$

$$\text{Percent yield} = \frac{1.85\ \text{g}}{2.20\ \text{g}} \times 100\ \%$$

$$\text{Percent yield} = 84.1\ \%$$

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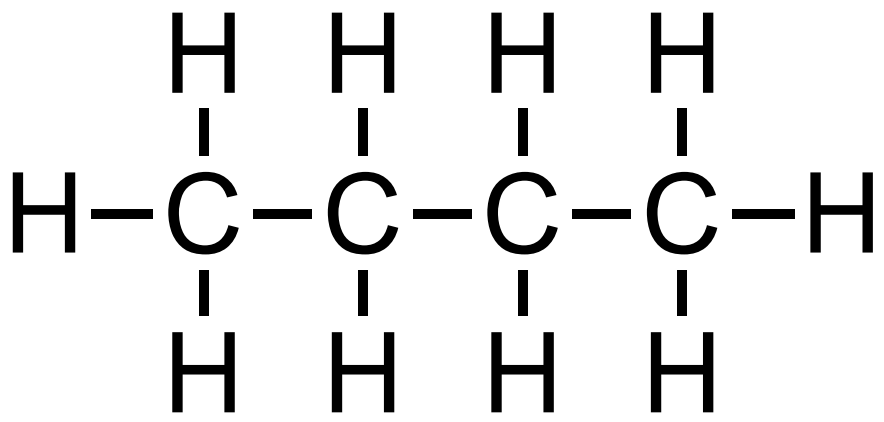
Empirical and molecular formula

Empirical and molecular formula

The molecular formula is the actual number of atoms in a compound.

The empirical formula is the lowest whole number ratio of atoms in a compound.

butane



molecular formula C_4H_{10}

empirical formula C_2H_5

Empirical and molecular formula

A compound has the empirical formula CH_2O and a molar mass of 180.18 g mol^{-1}. Determine its molecular formula.

$$12.01 + (2 \times 1.01) + 16.00 = 30.03$$

$$\frac{180.18}{30.03} = 6$$

$$CH_2O \times 6 = C_6H_{12}O_6$$

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Calculating % composition by mass

Calculating % composition by mass

Calculate the percentage composition by mass of Na, S and O in Na_2SO_4

$$\% \text{ composition by mass} = \frac{\text{total mass of element}}{\text{total mass of compound}} \times 100$$

Determine the molar mass of the compound:

$$(2 \times 22.99) + (32.07) + (4 \times 16.00) = 142.05 \text{ gmol}^{-1}$$

Calculating % composition by mass

$$\% \text{ Na by mass} = \frac{2 \times 22.99}{142.05} \times 100 = 32.37\%$$

$$\% \text{ S by mass} = \frac{32.07}{142.05} \times 100 = 22.58\%$$

$$\% \text{ O by mass} = \frac{4 \times 16.00}{142.05} \times 100 = 45.05\%$$

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Calculating empirical formula from % composition by mass

Calculating empirical formula

An organic compound A contains 62.0% by mass of carbon, 24.1% by mass of nitrogen, and 13.9% by mass of hydrogen.

C	H	N
$\frac{62.0}{12.01}$	$\frac{13.9}{1.01}$	$\frac{24.1}{14.01}$
5.16	13.8	1.72

Calculating empirical formula

5.16	13.8	1.72
1.72	1.72	1.72
3	8	1

C_3H_8N

Calculating empirical formula

The molar mass of A is 116 g mol^{-1}. Determine the molecular formula of A.

$$C_3H_8N$$

$$(3 \times 12.01) + (8 \times 1.01) + (1 \times 14.01) = 58.12$$

$$116 \div 58.12 = 2$$

$$2 \times C_3H_8N = C_6H_{16}N_2$$

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Percentage purity

Percentage purity

Percentage purity is the percentage of a pure compound in an impure sample.

$$\%\ \text{purity} = \frac{\text{mass of pure compound in sample}}{\text{total mass of impure sample}} \times 100$$

A 150.0 g sample of copper ore contains 87.3 g of pure copper. Calculate the percentage purity.

$$\%\ \text{purity} = \frac{87.3}{150.0} \times 100 = 58.2\ \%$$

Percentage purity

10.00 g of chalk (impure $CaCO_3$) was reacted with excess HCl.

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$$

2.13 dm^3 of CO_2 was produced at STP in the reaction.

$$n(CO_2) = 2.13 / 22.7 = 9.38 \times 10^{-2} \text{ mol}$$

$$m(CaCO_3) = 9.38 \times 10^{-2} \times 100.09 = 9.38 \text{ g}$$

$$\% \text{ purity} = (9.38 / 10.00) \times 100 = 93.8 \%$$

Percentage purity

5.00 g of an impure sample of hydrated ethanedioic acid, $(COOH)_2 \bullet 2H_2O$, was dissolved in water to make 1.00 dm^3 of solution.

25.0 cm^3 samples of this solution were titrated against a 0.100 mol dm^{-3} solution of sodium hydroxide using a suitable indicator.

The average value of the titre (NaOH) was 14.0 cm^3.

$$(COOH)_{2(aq)} + 2NaOH_{(aq)} \rightarrow (COONa)_{2(aq)} + 2H_2O_{(l)}$$

Percentage purity

$n(\text{NaOH}) = 0.100 \times (14.0 / 1000) = 1.40 \times 10^{-3}$ mol

$n((\text{COOH})_2 \bullet 2\text{H}_2\text{O}) = 1.40 \times 10^{-3} / 2 = 7.00 \times 10^{-4}$ mol

In 1 dm^3: $7.00 \times 10^{-4} \times (1000 / 25) = 2.80 \times 10^{-2}$ mol

$m((\text{COOH})_2 \bullet 2\text{H}_2\text{O}) = 2.80 \times 10^{-2} \times 126.08 = 3.53$ g

% purity = $(3.53 / 5.00) \times 100 = 70.6$ %

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Calculating empirical formula from combustion analysis

Empirical formula

Menthol is an organic compound composed of C, H and O atoms.

The complete combustion of 0.1005 g of menthol produces 0.2829 g of CO_2 and 0.1159 g of H_2O. Calculate the empirical formula of menthol.

$$C_xH_yO_z + \_O_2 \rightarrow \_CO_2 + \_H_2O$$

0.1005 g 0.2829 g 0.1159 g

Empirical formula

Calculate the mass of C in 0.2829 g of CO_2

$$\frac{12.01}{44.01} \times 0.2829 = 0.07720 \text{ g of C}$$

Convert to amount in mol (n)

$$n = \frac{0.07720}{12.01} = 6.428 \times 10^{-3} \text{ mol C}$$

Empirical formula

Calculate the mass of H in 0.1159 g of H_2O

$$\frac{2.02}{18.02} \times 0.1159 = 0.01299 \text{ g of H}$$

Convert to amount in mol (n)

$$n = \frac{0.01299}{1.01} = 0.01286 \text{ mol H}$$

Empirical formula

Calculate the mass of O in 0.1005 g of menthol

$$0.1005 - 0.07720 - 0.01299 = 0.01031 \text{ g O}$$

Convert to amount in mol (*n*)

$$n = \frac{0.01031}{16.00} = 6.444 \times 10^{-4} \text{ mol O}$$

Empirical formula

Divide each amount by the smallest to get the lowest whole number ratio:

6.428×10^{-3} mol C	0.01286 mol H	6.444×10^{-4} mol O
6.444×10^{-4}	6.444×10^{-4}	6.444×10^{-4}
10	20	1

Empirical formula: $C_{10}H_{20}O$

Empirical formula

The molar mass of menthol is 156.30 g mol^{-1}. Calculate the molecular formula of menthol.

Empirical formula: $C_{10}H_{20}O$

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Molar volume of a gas

Molar volume of a gas

The molar volume of a gas (V_m) is the volume occupied by one mole of an ideal gas.

$$\text{STP}: 273\ \text{K and } 1.00 \times 10^5\ \text{Pa}$$

$$V = \frac{1.00 \times 8.31 \times 273}{1.00 \times 10^5} = 0.0227\ \text{m}^3$$

$$V_m = 0.0227\ \text{m}^3\text{mol}^{-1} \text{ or } 22.7\ \text{dm}^3\ \text{mol}^{-1}$$

Molar volume of a gas

The molar volume of a gas (V_m) is the volume occupied by one mole of an ideal gas.

$$\text{STP}: 273\ \text{K and } 1.00 \times 10^5\ \text{Pa}$$

$$V_m\ (\text{m}^3\ \text{mol}^{-1}) = \frac{0.0227\ \text{m}^3}{1\ \text{mol}}$$

$$V_m = 0.0227\ \text{m}^3\ \text{mol}^{-1}\ \textit{or}\ 22.7\ \text{dm}^3\ \text{mol}^{-1}$$

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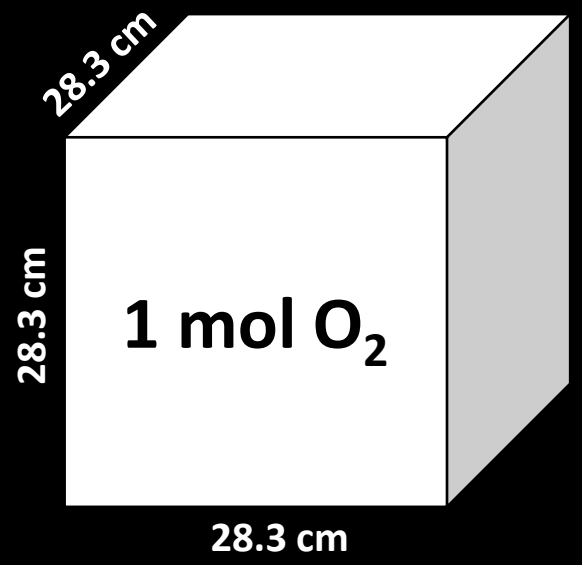
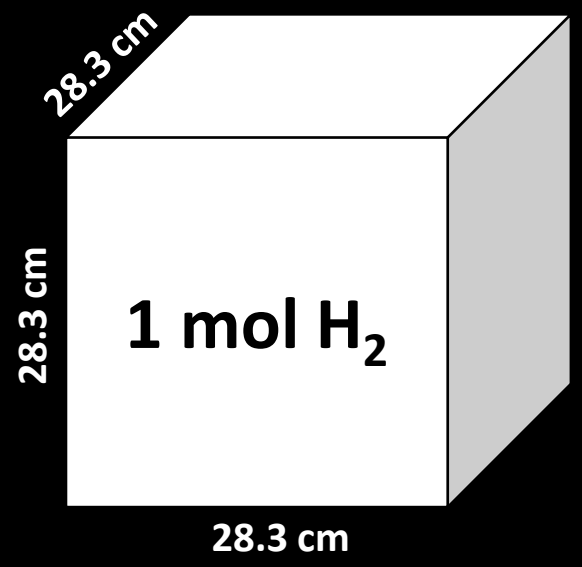
Molar volume of a gas

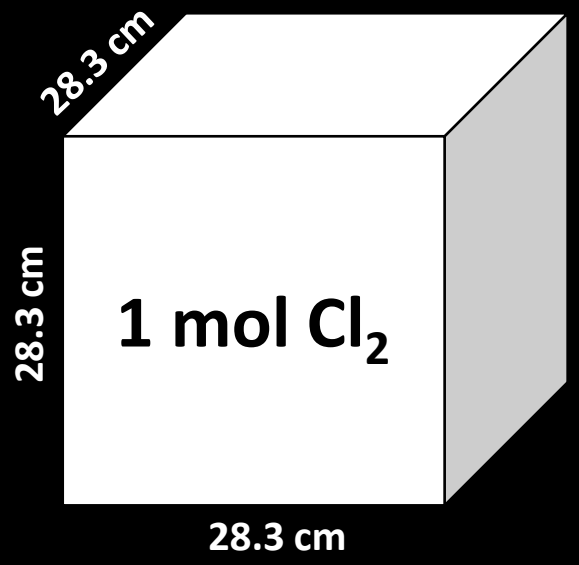
Under conditions of STP one mole of any gas occupies a volume of 0.0227 m^3 (22.7 dm^3).



22.7 dm^3



22.7 dm^3



22.7 dm^3

Molar volume of a gas

$$V\left(\mathrm{dm}^3\right) = n\,(\mathrm{mol}) \times V_m\left(22.7\ \mathrm{dm}^3\right)$$

$$V = n \times 22.7$$

$$n\,(\mathrm{mol}) = \frac{V\left(\mathrm{dm}^3\right)}{V_m\left(22.7\ \mathrm{dm}^3\right)}$$

$$n = \frac{V}{22.7}$$

Molar volume of a gas

Calculate the volume (in dm^3) occupied by 0.250 mol of N_2 at STP.

$$V = n \times 22.7$$

$$V = 0.250 \times 22.7$$

$$V = 5.68\ \mathrm{dm}^3$$

Molar volume of a gas

Calculate the volume (in cm^3) occupied by 0.00619 mol of CO_2 at STP.

$$V = n \times 22.7$$

$$V = 0.00619 \times 22.7$$

$$V = 0.141\ \mathrm{dm}^3$$

$$0.141\ \mathrm{dm}^3 \times \frac{1000\ \mathrm{cm}^3}{1\ \mathrm{dm}^3} = 141\ \mathrm{cm}^3$$

Molar volume of a gas

Calculate the amount (in mol) of N_2 in a 0.742 dm^3 sample.

$$n = \frac{V}{22.7}$$

$$n = \frac{0.742}{22.7}$$

$$n = 0.0327 \text{ mol}$$

Molar volume of a gas

Calculate the amount (in mol) of CH_4 in a 2.36 cm^3 sample.

$$2.36\ \mathrm{cm}^3 \times \frac{1\ \mathrm{dm}^3}{1000\ \mathrm{cm}^3} = 2.36 \times 10^{-3}\ \mathrm{dm}^3$$

$$n = \frac{V}{22.7} \qquad n = \frac{2.36 \times 10^{-3}}{22.7}$$

$$n = 1.04 \times 10^{-4}\ \mathrm{mol}$$

Molar volume of a gas

Determine the volume of H_2 (in cm^3) produced at STP when 2.00 g of Mg is reacted with excess $HCl_{(aq)}$.

$$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

$$n(\mathrm{Mg}) = \frac{2.00}{24.31} = 0.0823\ \mathrm{mol}$$

Ratio of Mg to H_2 is 1:1

0.0823 mol Mg will produce 0.0823 mol H_2

Molar volume of a gas

$$n(\mathrm{H_2}) = 0.0823\ \mathrm{mol}$$

$$V = n \times 22.7$$

$$V = 0.0823 \times 22.7$$

$$V = 1.87\ \mathrm{dm^3}$$

$$1.87\ \mathrm{dm^3} \times \frac{1000\ \mathrm{cm^3}}{1\ \mathrm{dm^3}} = 1.87 \times 10^3\ \mathrm{cm^3}$$

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Ideal gas equation

Ideal gas equation

$$V \propto \frac{1}{P} \qquad V \propto T \qquad V \propto n$$

$$V \propto \frac{nT}{P} \qquad V = R\left(\frac{nT}{P}\right)$$

$$PV = nRT$$

Ideal gas equation

$$PV = nRT$$

P is pressure (Pa)

V is volume (m^3)

n is amount (mol)

R is the gas constant ($8.31\ J\ K^{-1}\ mol^{-1}$)

T is temperature (K)

Ideal gas equation

$$n = \frac{PV}{RT} \qquad V = \frac{nRT}{P}$$

$$P = \frac{nRT}{V} \qquad T = \frac{PV}{nR}$$

Ideal gas equation

Unit conversions

Temperature in kelvin (K): $^{o}C + 273$

$25^{o}C = 298\ K$

Pressure in Pa: 1.00×10^{5} Pa = 100 kPa

$1\ cm^3 = 1 \times 10^{-3}\ dm^3 = 1 \times 10^{-6}\ m^3$

$1\ m^3 = 1 \times 10^{3}\ dm^3 = 1 \times 10^{6}\ cm^3$

Ideal gas equation

$\times 1000$ $\times 1000$

$$1\ \mathrm{m}^3 = 1 \times 10^3\ \mathrm{dm}^3 = 1 \times 10^6\ \mathrm{cm}^3$$

$\div 1000$ $\div 1000$

Ideal gas equation

Calculate the volume (in dm^3) occupied by 0.500 mol of gas at 1.50×10^5 Pa and 25.0 °C.

$$V = \frac{nRT}{P} \qquad V = \frac{0.500 \times 8.31 \times 298}{150000}$$

$$V = 8.25 \times 10^{-3}\ \mathrm{m}^3 = 8.25\ \mathrm{dm}^3$$

Ideal gas equation

Calculate the pressure (in kPa) of 0.200 mol of gas that occupies a volume of 10.0 dm^3 at 20.0 ^oC.

$$P = \frac{nRT}{V} \qquad P = \frac{0.200 \times 8.31 \times 293}{0.0100}$$

$$P = 4.87 \times 10^4 \text{ Pa} = 48.7 \text{ kPa}$$

Ideal gas equation

Calculate the amount (in mol) of gas that occupies a volume of 20.0 dm^3 at 50.0 °C and 85.0 kPa.

$$n = \frac{PV}{RT}$$

$$n = \frac{85000 \times 0.0200}{8.31 \times 323}$$

$$n = 0.633 \text{ mol}$$

Ideal gas equation

1.10 g of an unknown gas occupies a volume of 567 cm^3 at STP. Calculate the molar mass of the gas.

$$M = \frac{mRT}{PV} \qquad M = \frac{1.10 \times 8.31 \times 273}{1.00 \times 10^5 \times 5.67 \times 10^{-4}}$$

$$M = 44.0 \text{ g mol}^{-1}$$

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The gas laws part 1

The gas laws

$$P \propto \frac{1}{V}$$

$$V \propto T$$

$$P \propto T$$

$$V \propto n$$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

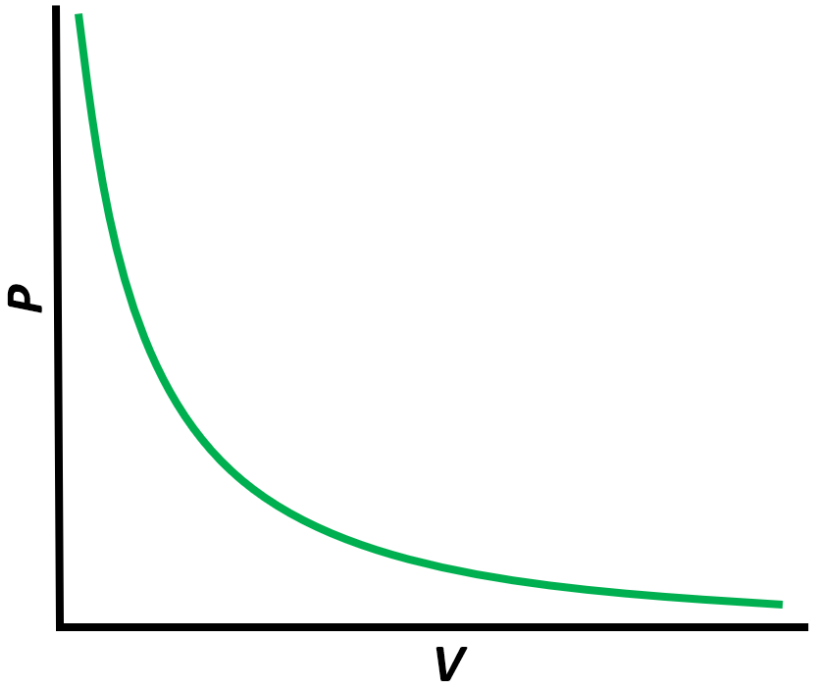
The gas laws

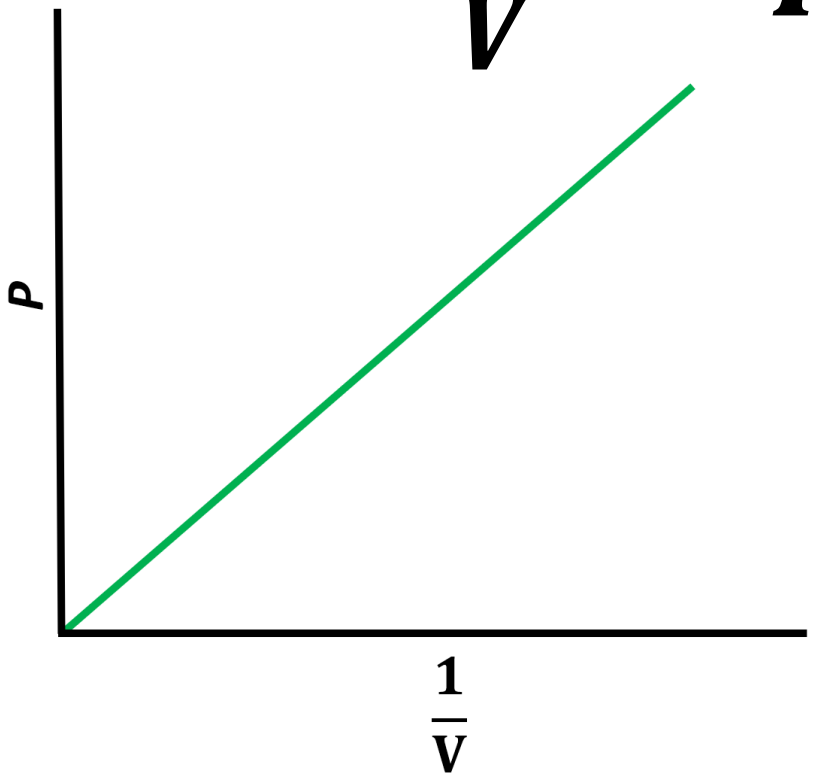
Boyle's law - the volume occupied by a gas is inversely proportional to its pressure (at constant *n* and *T*).

$$PV = k$$

$$P \propto \frac{1}{V}$$

$$P_1V_1 = P_2V_2$$



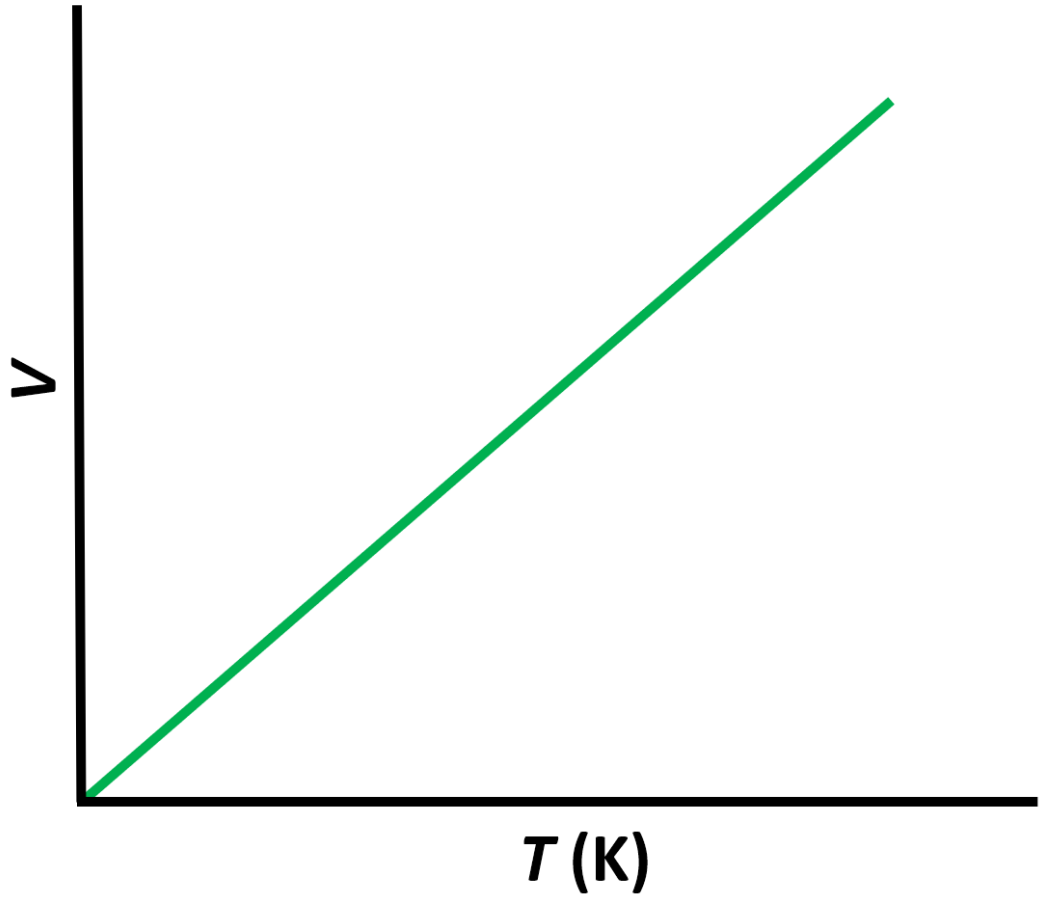


The gas laws

Charles's law - the volume occupied by a gas is directly proportional to its absolute temperature (at constant *n* and *P*).

$$V \propto T \qquad \frac{V}{T} = k$$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

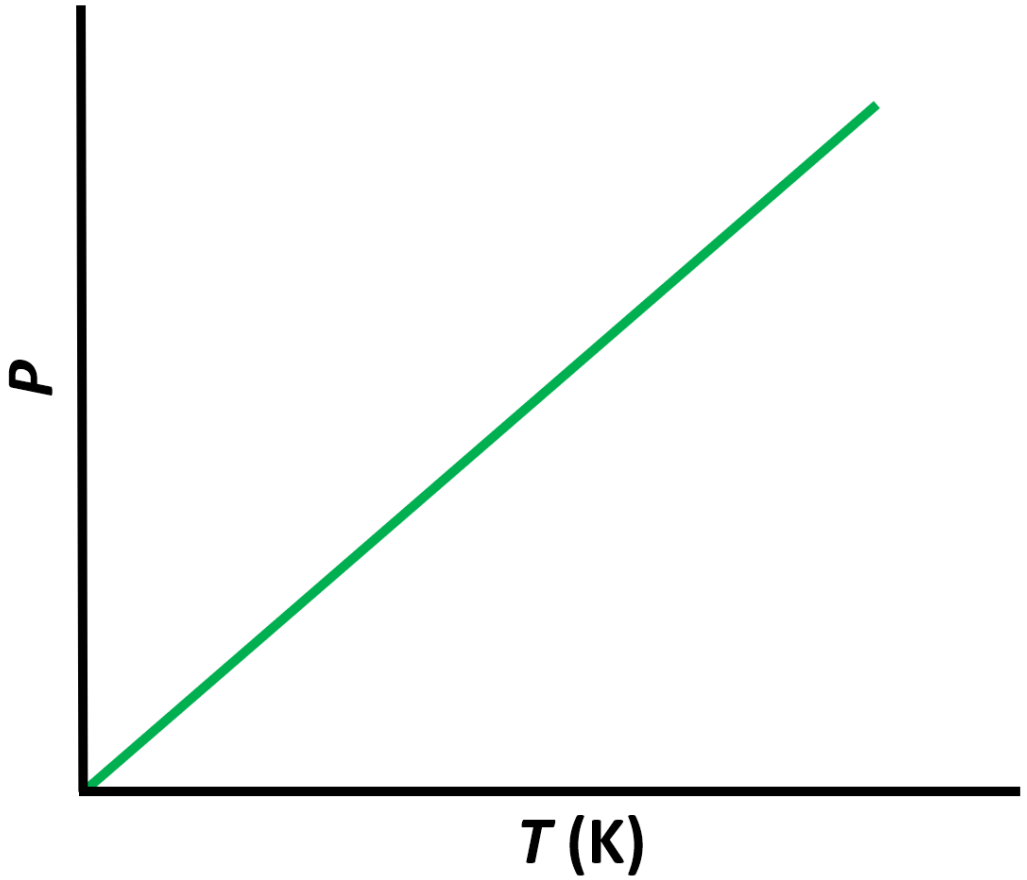


The gas laws

Gay Lussac's law - the pressure exerted by a gas is directly proportional to its absolute temperature (at constant *n* and *V*).

$$P \propto T \qquad \frac{P}{T} = k$$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

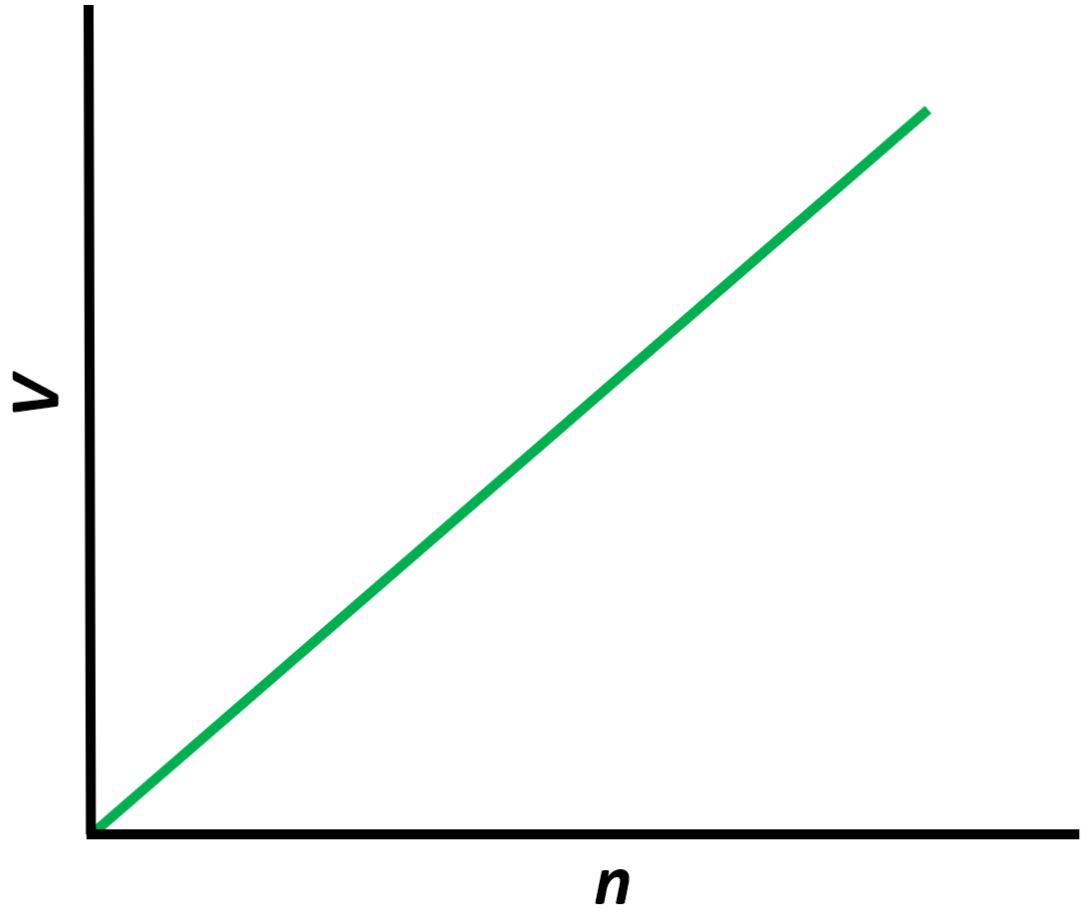


The gas laws

Avogadro's law - the volume occupied by a gas is directly proportional to the amount (in mol) of gas (at constant *P* and *T*).

$$V \propto n \qquad \frac{V}{n} = k$$

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$



The gas laws

The Combined gas law combines Boyle's law, Charles's law and Gay-Lussac's law.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad \frac{P_1V_1}{T_1} = k$$

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The gas laws part 2

The gas laws

$$P \propto \frac{1}{V}$$

$$V \propto T$$

$$P \propto T$$

$$V \propto n$$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

The gas laws

A sample of gas has a volume of 15.0 cm^3 at a pressure of 575 kPa. Assuming that temperature remains constant, what volume will the gas occupy at a pressure of 968 kPa?

$$P_1V_1 = P_2V_2$$

$$V_2 = \frac{P_1V_1}{P_2}$$

$$V_2 = \frac{575 \times 15.0}{968}$$

$$V_2 = 8.91\ \mathrm{cm}^3$$

The gas laws

A sample of gas has a volume of 32.0 dm^3 at a temperature of 256 K. Assuming that pressure remains constant, what volume will the gas occupy at a temperature of 391 K?

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$V_2 = \frac{V_1 T_2}{T_1}$$

$$V_2 = \frac{32.0 \times 391}{256}$$

$$V_2 = 48.9\ \mathrm{dm}^3$$

The gas laws

A sample of gas has a pressure of 73.9 kPa at a temperature of 347 K. Assuming that volume remains constant, what will be the pressure of the gas at a temperature of 602 K?

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$P_2 = \frac{P_1 T_2}{T_1}$$

$$P_2 = \frac{73.9 \times 602}{347}$$

$$P_2 = 128\ \text{kPa}$$

The gas laws

A sample contains 5.13 mol of gas with a volume of 1.28 m^3. Assuming that temperature and pressure remain constant, what volume will the gas occupy if 3.49 mol of gas are added?

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

$$V_2 = \frac{V_1 n_2}{n_1}$$

$$V_2 = \frac{1.28 \times 8.62}{5.13}$$

$$V_2 = 2.15\ \mathrm{m}^3$$

The gas laws

A sample of gas has a volume of 1.54 m^3 at a temperature of 447 K and a pressure of 12.0 kPa. If the temperature and pressure are changed to 561 K and 15.7 kPa respectively, what volume will the gas occupy?

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = \frac{1.54 \times 12.0 \times 561}{447 \times 15.7}$$

$$V_2 = \frac{V_1P_1T_2}{T_1P_2}$$

$$V_2 = 1.48\ \mathrm{m}^3$$

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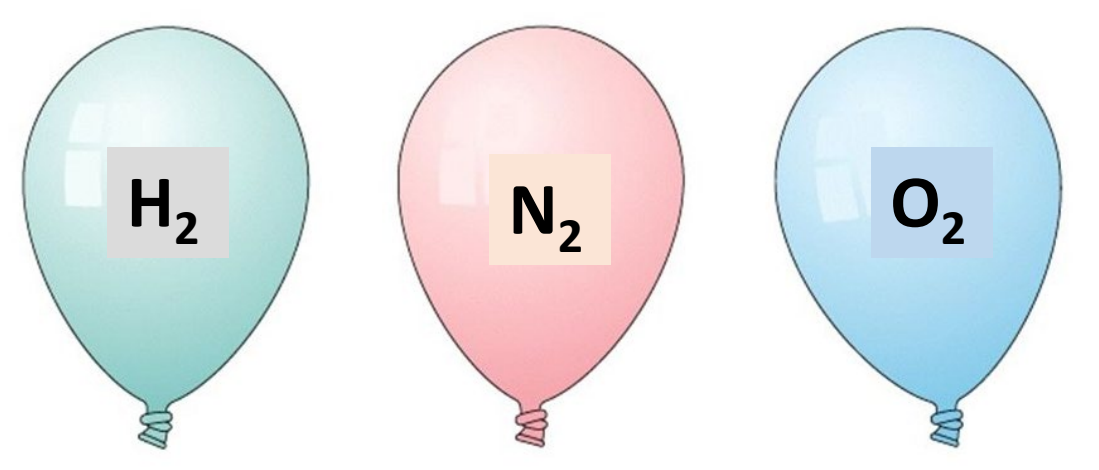
Avogadro's law

Avogadro's law

Equal volumes of gases at the same temperature and pressure contain the same number of particles.

STP
273 K
1.00×10^5 Pa



amount (mol)	1 mol H_2	1 mol N_2	1 mol O_2
volume (dm^3)	22.7	22.7	22.7
# of particles	6.02×10^{23}	6.02×10^{23}	6.02×10^{23}

Avogadro's law

40 cm^3 of CO reacts with 40 cm^3 of O_2. What volume of CO_2 is produced? What volume of the excess reactant remains?

$$2CO_{(g)} + O_{2(g)} \rightarrow 2CO_{2(g)}$$

limiting reactant excess reactant

2 vol	:	1 vol	:	2 vol
40 cm^3	:	40 cm^3	:	?

Avogadro's law

$$2CO_{(g)} + O_{2(g)} \rightarrow 2CO_{2(g)}$$

$40\ cm^3 : 40\ cm^3 : ?$

$0\ cm^3 : 20\ cm^3 : 40\ cm^3$

Avogadro's law

100 cm^3 of ethene (C_2H_4) is burned in 400 cm^3 of oxygen, producing carbon dioxide and liquid water. Calculate the volume of carbon dioxide produced and the volume of oxygen remaining.

$$C_2H_{4(g)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 2H_2O_{(l)}$$

limiting reactant excess reactant

1 vol : 3 vol : 2 vol

100 cm^3 : 400 cm^3 : ?

Avogadro's law

$$C_2H_{4(g)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 2H_2O_{(l)}$$

$100\ cm^3 : 400\ cm^3 : \ ?$

$0\ cm^3 : 100\ cm^3 : 200\ cm^3$

25	16		6	2	25
Mn	S	J	C	He	Mn
Manganese	Sulfur		Carbon	Helium	Manganese
54.938045	32.065		12.0107	4.002602	54.938045

Deviation from ideal gas behaviour

Ideal gases vs real gases

An ideal gas is a hypothetical gas that obeys the gas laws and the kinetic-molecular theory.

- Particles of an ideal gas are in constant, random, straight-line motion.
- Collisions between particles of an ideal gas are elastic; total kinetic energy is conserved.
- The volume occupied by the particles of an ideal gas is negligible relative to the volume of the container.
- There are no intermolecular forces acting between particles of an ideal gas.
- The average kinetic energy of the particles of an ideal gas is directly proportional to the absolute temperature in kelvin.

Ideal gases vs real gases

A real gas is a gas that deviates from ideal gas behaviour.

- Real gases have a finite, measurable volume.
- Real gases have intermolecular forces that act between the particles.

Real gases exhibit nearly ideal behaviour at relatively high temperatures and low pressures.

They deviate the most from ideal behaviour at low temperatures and high pressures.

Ideal gases vs real gases

Gas	Molar volume ($dm^3\ mol^{-1}$)
He	22.435
H_2	22.432
Ne	22.422
Ideal gas	22.414
Ar	22.397
N_2	22.396
O_2	22.390

Ideal gases vs real gases

For one mole of an ideal gas, the product of PV/RT is equal to one (regardless of the temperature or pressure).

$$n = \frac{1.00 \times 10^5 \text{ Pa} \times 0.0227 \text{ m}^3}{8.31 \text{ J K}^{-1} \text{ mol}^{-1} \times 273 \text{ K}} = 1.00 \text{ mol}$$

Real gases exhibit nearly ideal behaviour at relatively high temperatures and low pressures.

Real gases deviate the most from ideal gas behavior at high pressures and low temperatures.

Ideal gases vs real gases

For one mole of an ideal gas, the product of PV/RT is always equal to one.

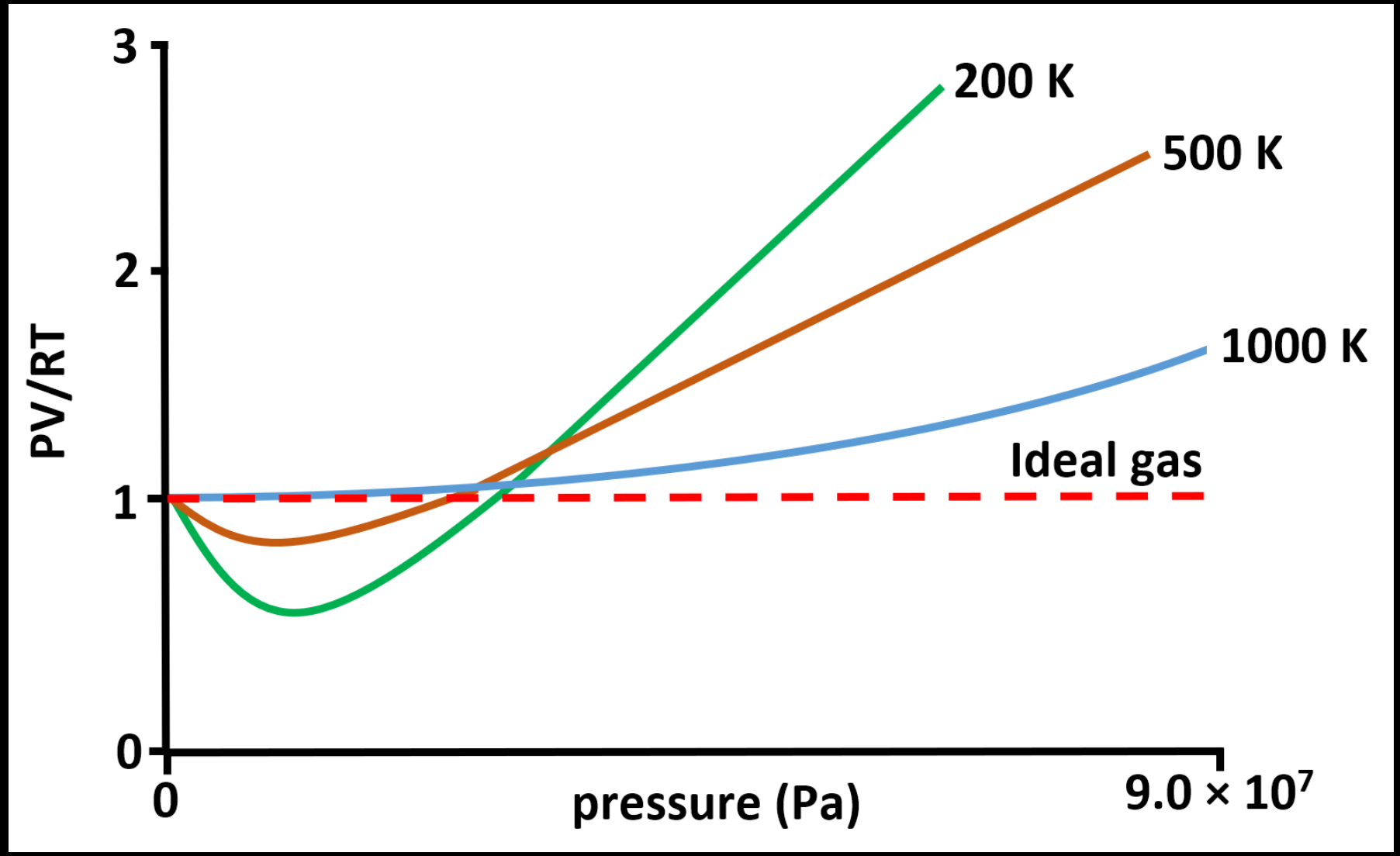
$$n = \frac{PV}{RT}$$

$$n = \frac{1.00 \times 10^5 \text{ Pa} \times 0.0227 \text{ m}^3}{8.31 \text{ J K}^{-1} \text{ mol}^{-1} \times 273 \text{ K}} = 1.00 \text{ mol}$$

For real gases the product of PV/RT ≠ 1.

Ideal gases vs real gases

Deviation of nitrogen gas from ideal gas behavior.

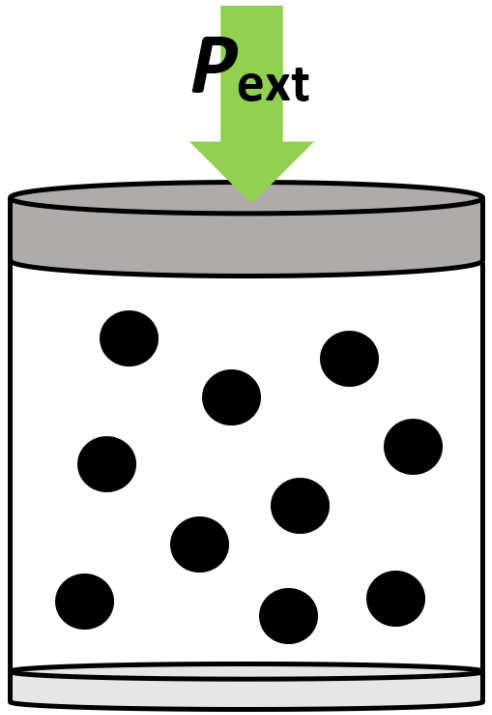


$$\frac{PV}{RT} > 1$$

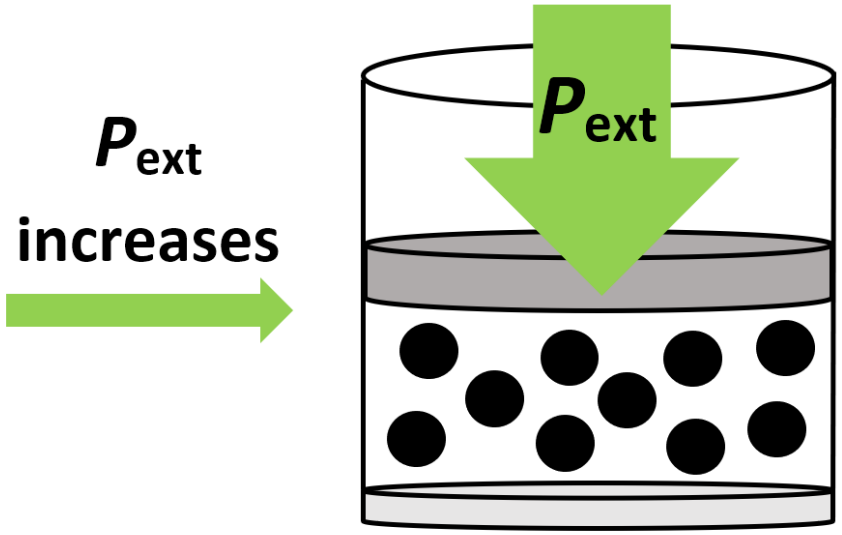
$$\frac{PV}{RT} < 1$$

Ideal gases vs real gases

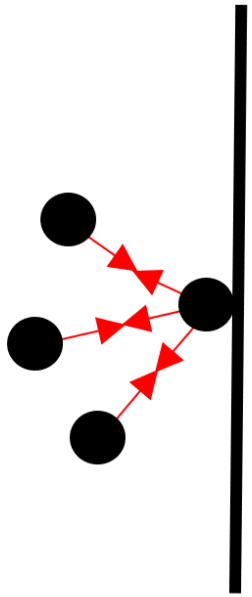
At moderately high pressures, the values of PV/RT are less than one, mainly because of the effects of intermolecular forces.



Lower P_{ext}; particles are too far apart for intermolecular forces to act



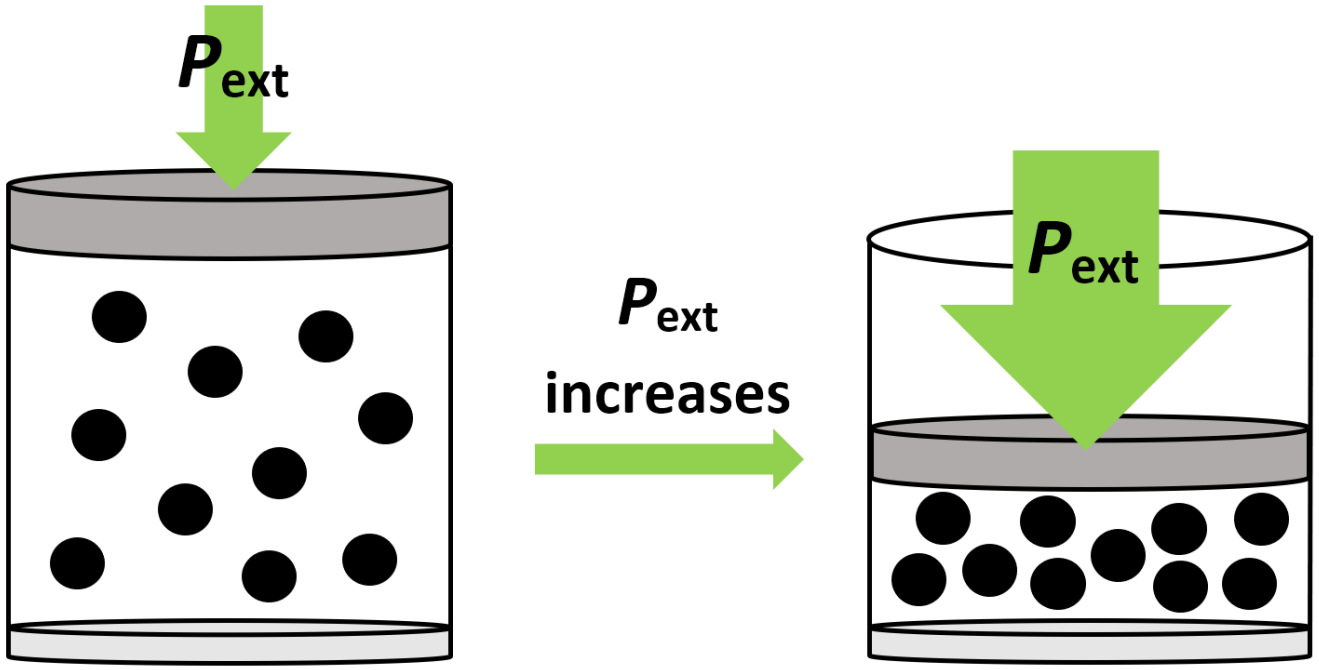
Moderately high P_{ext}; particles are now close enough for intermolecular forces to act



Intermolecular attractions reduce the force of the collisions with the container wall which results in a lower pressure

Ideal gases vs real gases

At very high pressures, the values of PV/RT are greater than one, mainly because of the effects of molecular volume.

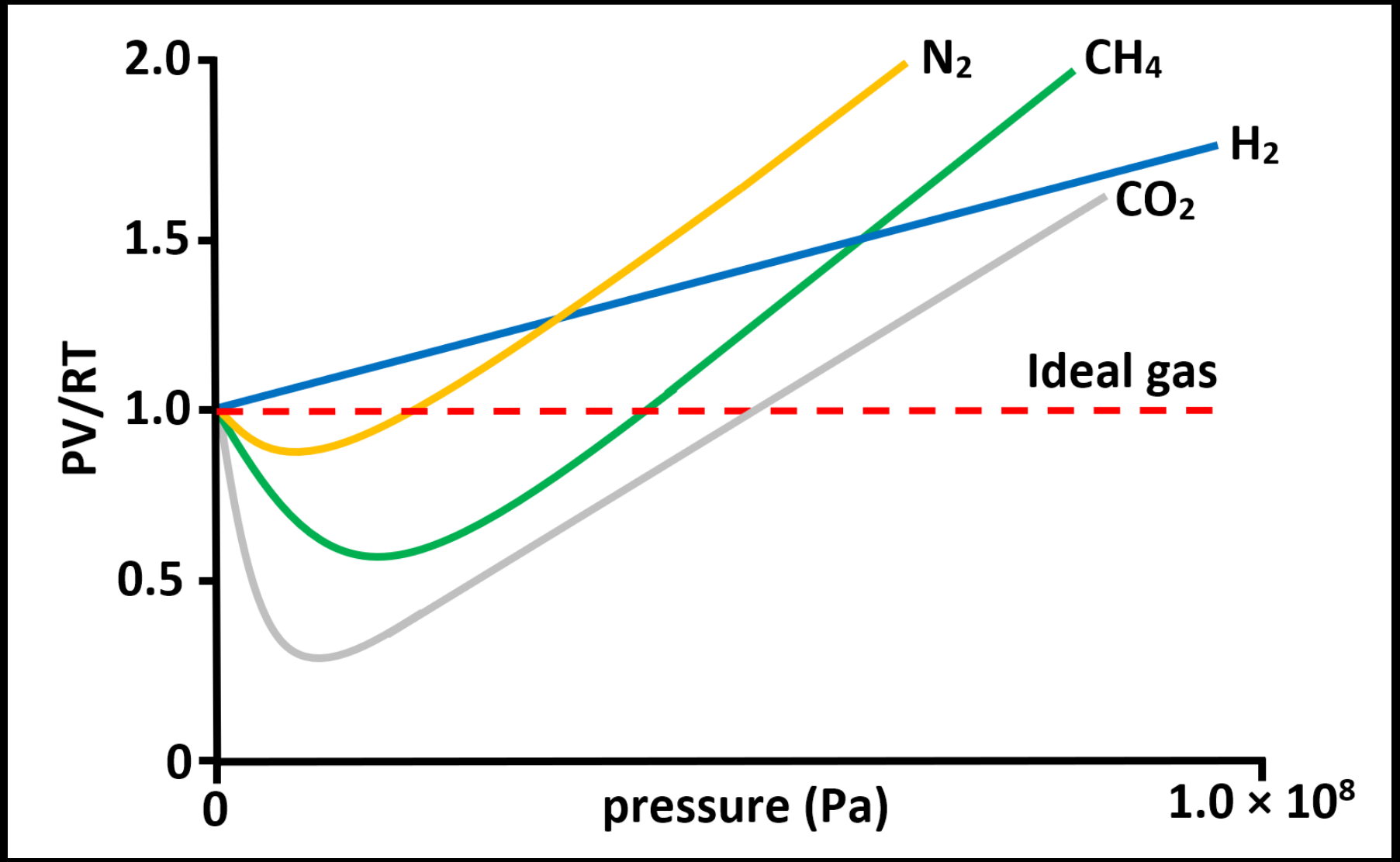


Lower P_{ext}; the volume occupied by the gas particles is negligible compared to the volume of the container

Very high P_{ext}; the volume occupied by the gas particles becomes significant

Ideal gases vs real gases

Deviation of different gases from ideal gas behaviour.



$$\frac{PV}{RT} > 1$$

$$\frac{PV}{RT} < 1$$

Ideal gases vs real gases

Ideal gases	Real gases
Ideal gases behave ideally at all temperatures and pressures	Real gases deviate the most from ideal behaviour at low temperatures and high pressures
The volume occupied by an ideal gas is assumed to be negligible	Real gases have a finite, measurable volume
Ideal gases have no intermolecular forces acting between the particles	Real gases have intermolecular forces acting between their particles
Ideal gases obey the ideal gas law $$PV = nRT$$	Real gases obey the van der Waals equation $$P = \frac{RT}{V-b} - \frac{a}{V^2}$$

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Concentration of solutions

Concentration of solutions

Concentration can be measured in mol dm^{-3}, g dm^{-3}, or ppm.

$$c\ (\mathrm{mol\ dm^{-3}}) = \frac{\text{amount of solute (mol)}}{\text{volume of solution (dm}^3\text{)}}$$

$$c\ (\mathrm{g\ dm^{-3}}) = \frac{\text{mass of solute (g)}}{\text{volume of solution (dm}^3\text{)}}$$

$$(\mathrm{ppm}) = \frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \times 10^6$$

Concentration of solutions

50.0 g of NaCl is dissolved in 100 cm^3 of distilled water and made up to 500 cm^3 in a volumetric flask. Calculate the concentration of the solution in g dm^{-3} and mol dm^{-3}.

Concentration of solutions

$$n = \frac{m}{M} = \frac{50.0}{58.44} = 0.856 \text{ mol NaCl}$$

$$c = \frac{n}{V} = \frac{0.856}{(500.0 \div 1000)}$$

$$c = 1.71 \text{ mol dm}^{-3}$$

Concentration of solutions

Calculate the mass (in g) of sodium carbonate (Na_2CO_3) needed to make a 250.0 cm^3 solution with a concentration of 0.500 mol dm^{-3}

$$n = cV$$

$$n = 0.500 \times (250.0 \div 1000)$$

$$n = 0.125 \text{ mol}$$

Concentration of solutions

$$m = nM$$

$$m = 0.125 \times 105.99$$

$$m = 13.2\ \text{g}\ Na_2CO_3$$

Concentration of solutions

An aqueous solution has 0.0070 g of oxygen dissolved in 1.00 dm^3 of water. Calculate the concentration in ppm.

$$\text{density of water} = 1\ \text{gcm}^{-3}$$

$$\text{ppm} = \frac{0.0070}{(1000 + 0.0070)} \times 10^6$$

$$= 7.0\ \text{ppm}$$

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Thermometric titration

Thermometric titration

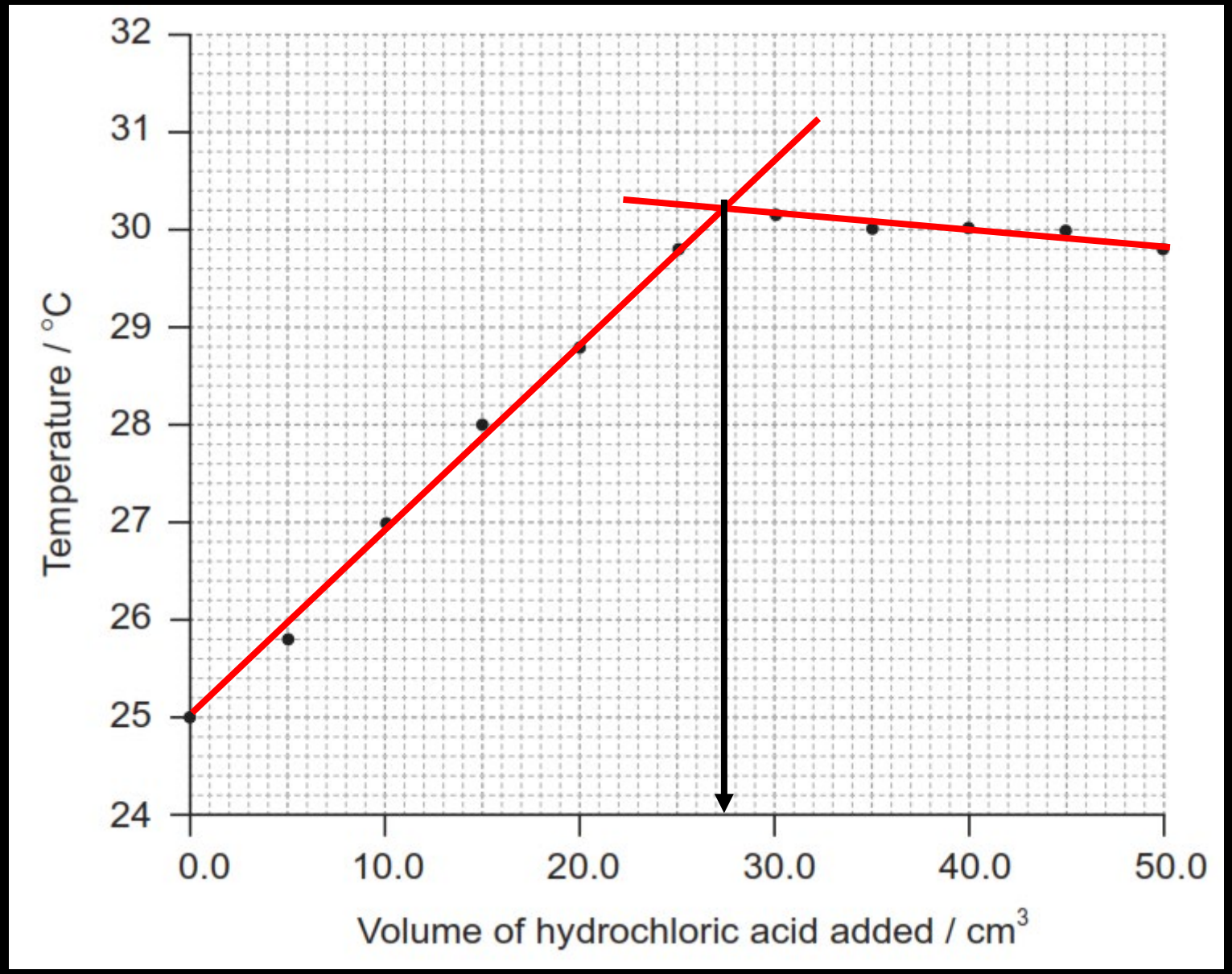
A thermometric titration can be carried out to determine the concentration of a solution.

An acid of unknown concentration is added to a base of known volume and concentration.

The temperature of the mixture is recorded as the acid is added to the base and the results are plotted on a graph.

The maximum temperature reached signifies the end-point of the titration.

Thermometric titration



50.0 cm^3 of HCl was added to 25.0 cm^3 of 1.00 mol dm^{-3} NaOH.

Volume of HCl required to neutralise the base = 27.0 cm^3

Thermometric titration

Calculate the concentration of the HCl, in mol dm^{-3}

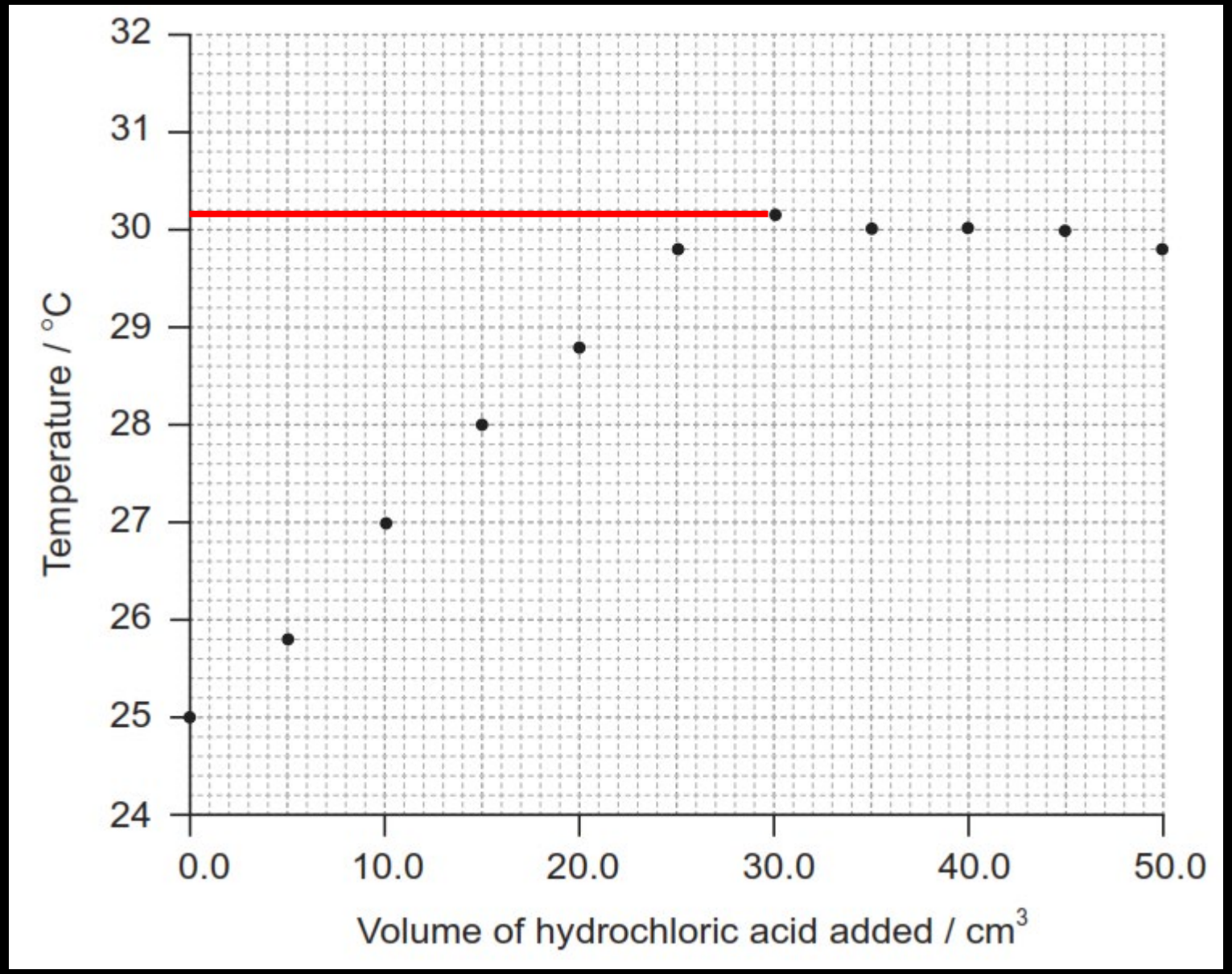
$$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

$$n\ (NaOH) = 1.00 \times (25.0 / 1000) = 0.0250\ mol$$

Ratio of NaOH : HCl = 1:1

$$C = 0.0250 \div (27.0 / 1000) = 0.926\ mol\ dm^{-3}$$

Thermometric titration



Maximum temperature reached = 30.2 °C

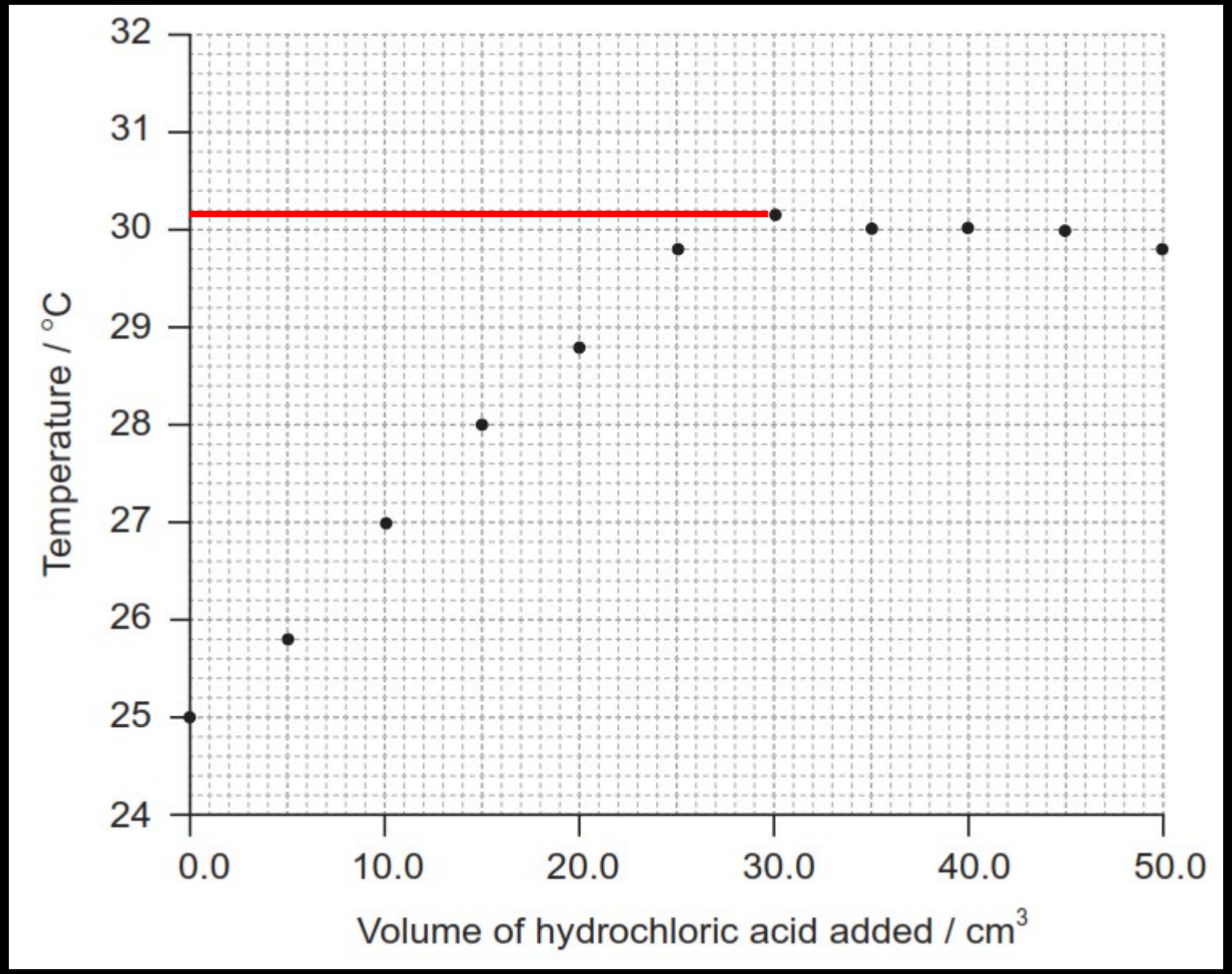
$\Delta T = 30.2 - 25.0$

$\Delta T = 5.20$ °C

ΔT assuming no heat loss

$30.6 - 25.0 = 5.60$ °C

Thermometric titration



Maximum temperature reached = 30.2 °C

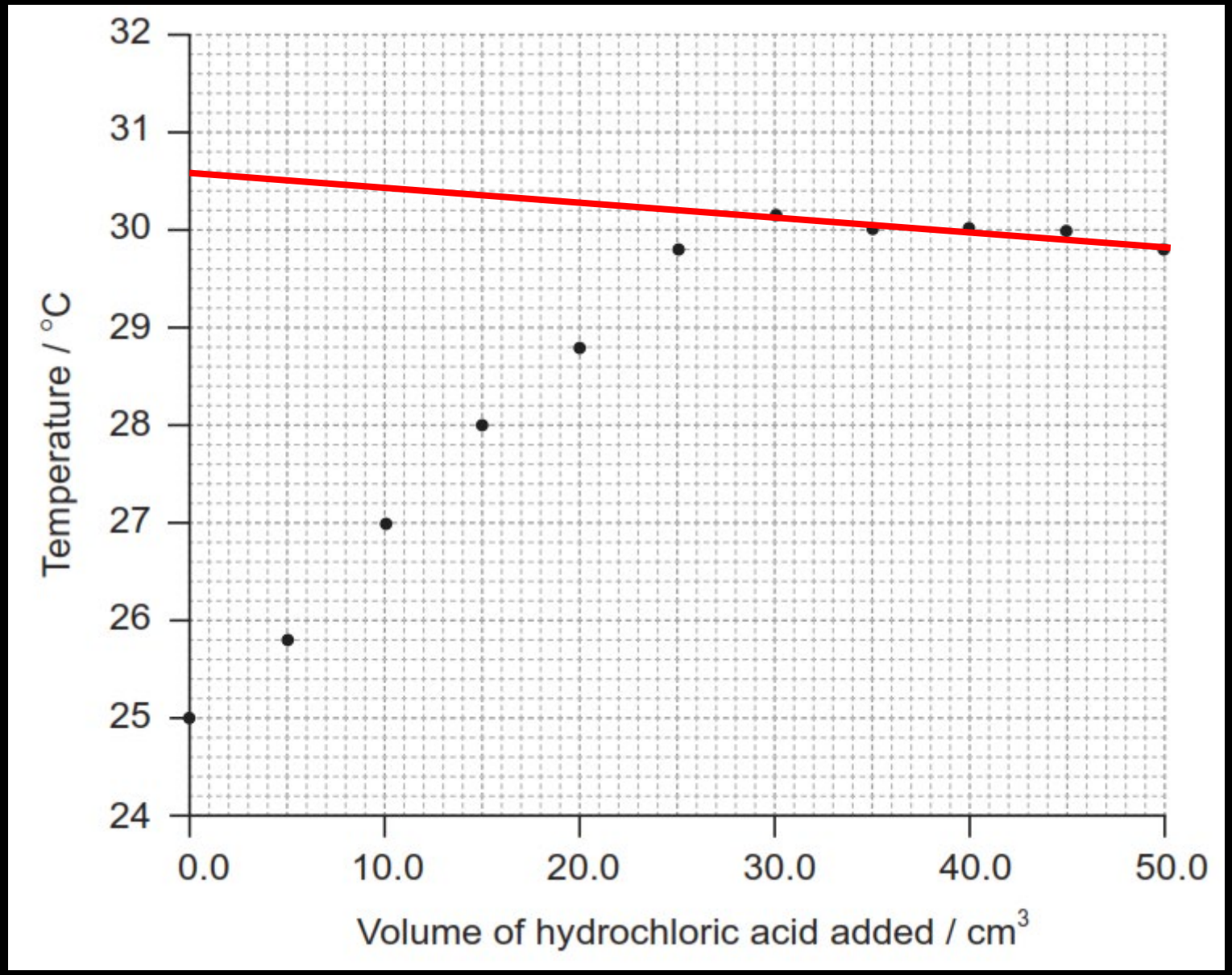
$\Delta T = 30.2 - 25.0$

$\Delta T = 5.20$ °C

ΔT assuming no heat loss

$30.6 - 25.0 = 5.60$ °C

Thermometric titration



Maximum temperature reached = 30.2 °C

ΔT = 30.2 – 25.0

ΔT = 5.20 °C

ΔT assuming no heat loss

30.6 – 25.0 = 5.60 °C

Thermometric titration

Calculate the heat change (q) for the reaction.

$$q = mc\Delta T$$

$$q = (25.0 + 27.0) \times 4.18 \times 5.60$$

$$q = 1220 \text{ J}$$

Thermometric titration

Calculate the enthalpy change (ΔH) in kJ mol^{-1} for the reaction.

$$\Delta H = q/n$$

$$\Delta H = 1220/0.0250$$

$$\Delta H = -48.8 \text{ kJ mol}^{-1}$$

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Calculating empirical formula from experimental data

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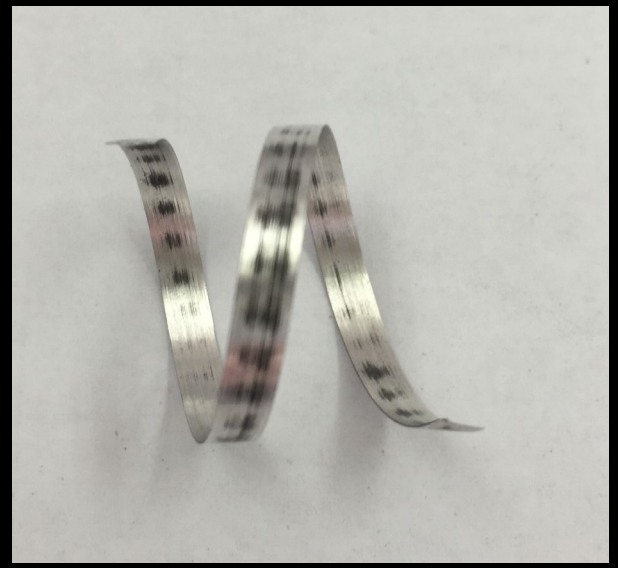
Calculating empirical formula



mass of crucible and lid

mass of magnesium, crucible and lid

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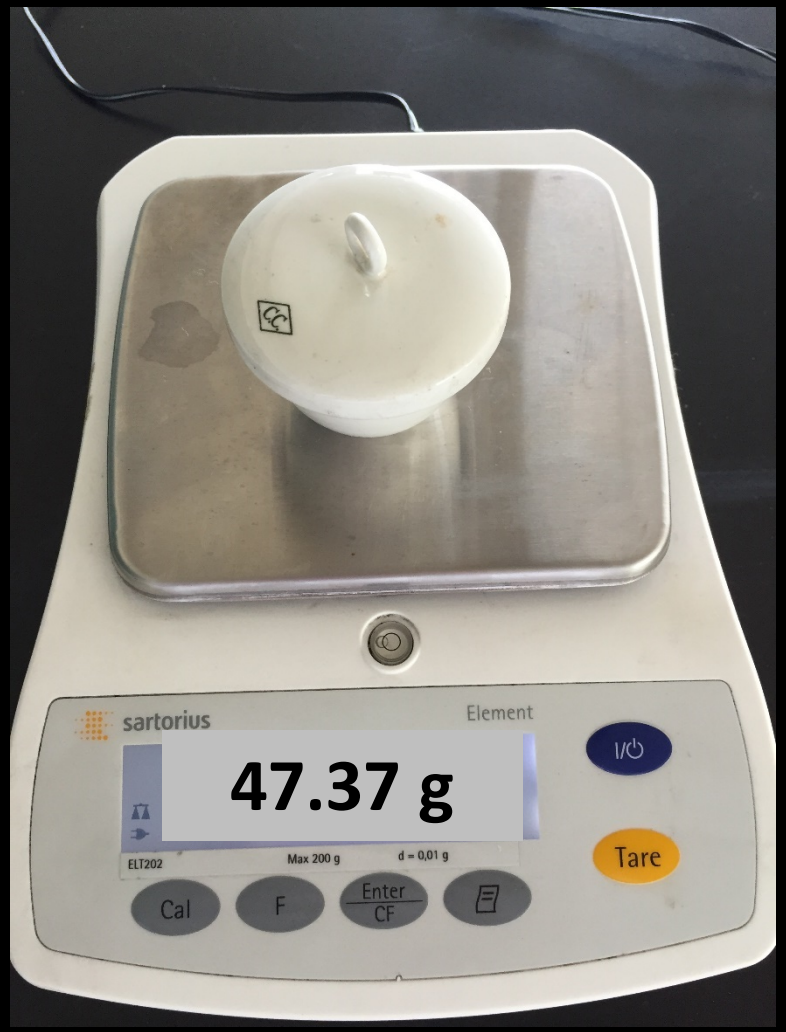
Calculating empirical formula

mass of contents, crucible and lid



Calculating empirical formula

Mass of crucible and lid ±0.01 g	40.19
Mass of magnesium, crucible and lid ±0.01 g	40.30
Mass of contents, crucible and lid ±0.01 g	40.37

Mass of Mg = mass of magnesium, crucible and lid – mass of crucible and lid

40.30 – 40.19 = 0.11g

Mass of O = mass of contents, crucible and lid – mass of magnesium, crucible and lid

40.37 – 40.30 = 0.07g

Calculating empirical formula

$$(n)\text{Mg} = \frac{0.11}{24.31} = 4.52 \times 10^{-3}\ \text{mol}$$

$$(n)\text{O} = \frac{0.07}{16.00} = 4.38 \times 10^{-3}\ \text{mol}$$

Calculating empirical formula

Mg	O
4.52×10^{-3}	4.38×10^{-3}
4.38×10^{-3}	4.38×10^{-3}
1.03	1

Empirical formula: MgO

Calculating empirical formula

In this experiment it is quite common not to get a 1:1 ratio of magnesium to oxygen.

Possible errors include:

- The magnesium was not pure.
- The product was not only magnesium oxide.
- The product (magnesium oxide) was lost when the lid was removed.

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Determining the molar mass of a gas experimentally

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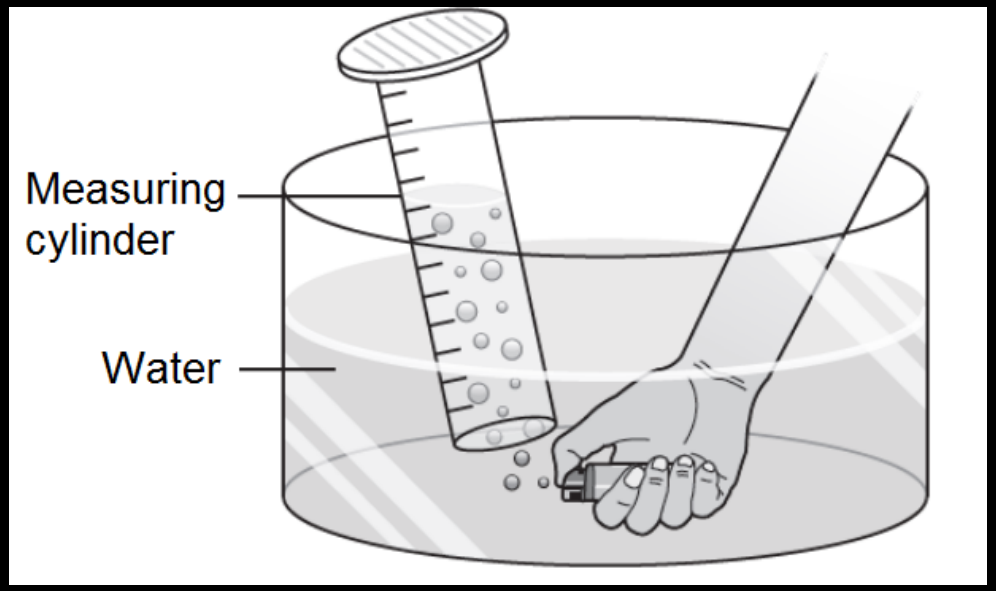
Tutorials for IB Chemistry

Calculating molar mass

The molar mass of a gas (such as butane, C_4H_{10}) can be determined experimentally.



Change in mass of lighter.
The volume of gas collected.
The pressure and the temperature.

$$M = \frac{mRT}{PV}$$

Calculating molar mass

Change in mass of lighter = 0.230 g

Volume of gas collected = 100.0 cm^3 = 1.000×10^{-4} m^3

Temperature = 22.0 °C = 295 K

Atmospheric pressure = 100717 Pa

Vapour pressure of water at 22.0 °C = 2634 Pa

$P_{(butane)}$ = 100717 – 2634 = 98083 Pa

R = 8.314 J K^{-1} mol^{-1}

$$M = \frac{0.230 \times 8.314 \times 295}{98083 \times 1.000 \times 10^{-4}}$$

M = 57.5 g mol^{-1}

Calculating molar mass

$$M\ C_4H_{10} = (4 \times 12.01) + (10 \times 1.01) = 58.14\ g\ mol^{-1}$$

$$\text{Percentage error} = \frac{57.5 - 58.14}{58.14} \times 100 = -1\ \%$$

Molar mass is lower than actual	Molar mass is higher than actual
Lighter is not dried completely before recording mass	Some bubbles of gas escaped from the cylinder
Pressure of gas not equalised with pressure of room	
Lighter may contain mixture of propane and butane	

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Water of crystallisation

Water of crystallisation

The water of crystallisation is the fixed number of water molecules present in one formula unit of a salt.

Hydrated salt
copper(II) sulfate
pentahydrate $CuSO_4 \cdot 5H_2O$

Anhydrous salt
copper(II) sulfate
$CuSO_4$

Water of crystallisation

Upon heating, the blue copper(II) sulfate pentahydrate decomposes forming the white anhydrous salt and water vapour.

$$CuSO_4 \bullet 5H_2O_{(s)} \rightarrow CuSO_{4(s)} + 5H_2O_{(g)}$$

Care must be taken as over-heating of the anhydrous salt can cause further decomposition to take place.

Water of crystallisation

1. Measure the mass of an empty crucible and lid.
2. Add a known mass of the sample and record the mass.
3. Heat the crucible for five minutes, holding the lid at an angle so the gas can escape.
4. After cooling, reweigh the crucible, lid and contents.
5. Repeat steps 3 and 4 until the mass remains constant (heating to constant mass).

Water of crystallisation

Determine the water of crystallisation for the hydrated salt $BaCl_2 \bullet xH_2O$ from the following data.

Mass of crucible and lid (g)	21.50
Mass of crucible, lid and sample (g)	26.50
Mass of crucible, lid and sample after 1st heating (g)	25.76
Mass of crucible, lid and sample after 2nd heating (g)	25.76

Mass of hydrated sample ($BaCl_2 \bullet xH_2O$) = 26.50 – 21.50 = 5.00 g

Mass of water driven off = 26.50 – 25.76 = 0.74 g

Mass of anhydrous $BaCl_2$ = 5.00 – 0.74 = **4.26 g**

Water of crystallisation

n ($BaCl_2$) = **4.26 / 208.23** = 0.02 mol

n (H_2O) = 0.74 ÷ 18.02 = 0.04 mol

$BaCl_2$: 0.02 ÷ 0.02 = 1

H_2O : 0.04 ÷ 0.02 = 2

$BaCl_2 \cdot 2H_2O$

Water of crystallisation

Assumptions made in the experiment:

All mass lost is due to the loss of water.

All water of crystallisation is driven off.

Crucible does not absorb water.

Anhydrous $BaCl_2$ does not decompose further.

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Back titration

Back titration

A back titration (indirect titration) has two stages:

- One reactant of unknown concentration is reacted with an excess reactant of known concentration.
- The amount of excess reactant is determined by a direct titration.

Back titrations are often used when one reactant is insoluble in water, for example salts such as calcium carbonate.

A common experiment is to determine the % calcium carbonate in an egg shell.

Back titration

A student carried out an experiment to determine the % $CaCO_3$ in an egg shell.

A 0.400 g sample of egg shell was crushed up and reacted with 25.00 cm^3 of 0.500 mol dm^{-3} HCl.

$$CaCO_{3\,(s)} + 2HCl_{\,(aq)} \rightarrow CaCl_{2\,(aq)} + H_2O_{\,(l)} + CO_{2\,(g)}$$

The excess acid was titrated with 0.500 mol dm^{-3} NaOH.

The average titre was 12.10 cm^3.

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

Back titration

Determine the amount (in mol) of HCl in 25.00 cm^3 of 0.500 mol dm^{-3} HCl.

$$n(HCl) = CV = 0.500 \times (25.00 / 1000) = 0.0125 \text{ mol}$$

Determine the amount (in mol) of excess HCl.

$$n(NaOH) = CV = 0.500 \times (12.10 / 1000) = 6.05 \times 10^{-3} \text{ mol}$$

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

$$n(HCl_{excess}) = 6.05 \times 10^{-3} \text{ mol}$$

Back titration

Determine the amount (in mol) of HCl that reacted with the $CaCO_3$ in the egg shell.

$$n(HCl_{reacted}) = n(HCl_{initial}) - n(HCl_{excess})$$

$$n(HCl_{reacted}) = 0.0125 - 6.05 \times 10^{-3} = 6.45 \times 10^{-3} \text{ mol}$$

Determine the amount (in mol) of $CaCO_3$ in the egg shell.

$$CaCO_{3\,(s)} + 2HCl_{\,(aq)} \rightarrow CaCl_{2\,(aq)} + H_2O_{\,(l)} + CO_{2\,(g)}$$

The ratio of $CaCO_3$ to HCl is 1:2.

Back titration

Divide the amount of HCl that reacted with the $CaCO_3$ by two.

$$n(CaCO_3) = 6.45 \times 10^{-3}\ \text{mol} / 2 = 3.23 \times 10^{-3}\ \text{mol}$$

Convert from amount (in mol) to mass (in g).

$$m = nM$$

$$m(CaCO_3) = 3.23 \times 10^{-3} \times 100.09$$

$$m(CaCO_3) = 0.323\ \text{g}$$

Back titration

Determine the percentage of $CaCO_3$ in the egg shell.

(mass of $CaCO_3$ / total mass of egg shell) × 100

(0.323 / 0.400) × 100 = 80.8 %

Possible sources of error:

- Not all the $CaCO_3$ in the egg shell reacted with the HCl.
- The egg shell was not dried for long enough.
- Not all of the membrane was removed.