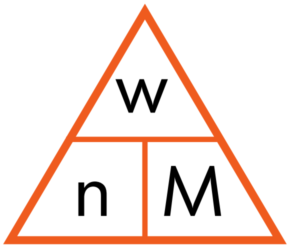
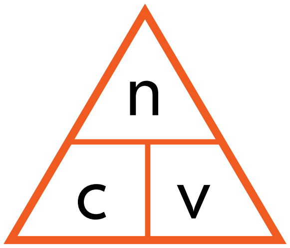
YEAR 1 CHEMISTRY EQUATIONS (AQA)

MOLE CALCULATIONS

number of moles = mass / molar mass (g) ($g.mol^{-1}$)

number of moles = concentration x volume ($mol.dm^{-3}$) (dm^3)

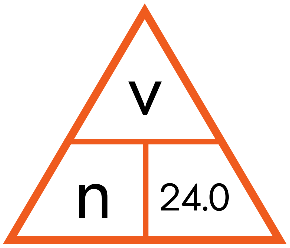




Avogadro's Constant = 6.02×10^{23} atoms or molecules = **1 mole**

MOLAR GAS CONSTANT

1 mole of ANY gas occupies 24.0 dm^3 at room temperature & pressure



IDEAL GAS EQUATION

P = Pressure (pa)
V = volume (m^3)
n = no. of moles
R = Gas Constant (8.314 $J.K^{-1}.mol^{-1}$)
T = Temperature (K)

$$PV = nRT$$

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

For changes in conditions:

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

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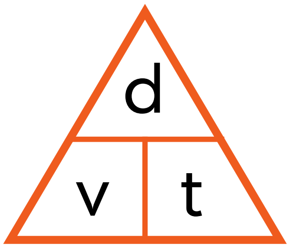
MASS SPECTROSCOPY

K.E. = Kinetic Energy (J)   m = mass of ONE ion (Kg)   v = velocity ($m.s^{-1}$)

$$K.E. = \tfrac{1}{2}\, m.v^2 \qquad v = \frac{\sqrt{2K.E.}}{m} \qquad m = \frac{\sqrt{2K.E.}}{v}$$

d = distance **or** length of flight tube (m)   t = time (s)   v = velocity ($m.s^{-1}$)



$$\text{Relative Atomic Mass (Ar)} = \frac{(\text{mass isotope 1 x abundance}) + (\text{mass isotope 2 x abundance}) + \ldots}{\sum \text{abundance}}$$

OTHER EQUATIONS

$$\text{\% by mass} = \frac{\text{mass of element in 1 mole}}{Mr}$$

$$\text{Empirical formula} = \frac{M1}{Mr1} : \frac{M2}{Mr2} : \frac{M3}{Mr2}$$

Where M1, M2 etc is the mass or % composition of element 1, 2 etc

then divide each by the smallest number to give empirical formula

$$\text{\% Atom Economy} = \frac{\text{mass of desired product}}{\text{total mass of all products}} \times 100$$

You can use mass or number of moles here!

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

You can replace masses with Mr values here too!

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ENTHALPY

Q = energy transferred (J)

m = mass of **solution** (g)

c = specific heat capacity ($J.K^{-1}.mol^{-1}$)

$\triangle T$ = **change** in temperature (°C **or** K)

$$\mathbf{Q = m.c.\triangle T}$$

$$\underset{(J.mol^{-1})}{\triangle H} = \frac{Q}{n}$$

Don't forget to add a sign for $\triangle H$!

Divide by 1000 for $kJ.mol^{-1}$

$$\underset{(kJ.mol^{-1})}{\triangle H \text{ reaction}} = \underset{(kJ.mol^{-1})}{\sum \text{reactant mean bond enthalpies}} - \underset{(kJ.mol^{-1})}{\sum \text{product mean bond enthalpies}}$$

EQUILIBRIA

$$aA + bB \rightleftharpoons cC + dD$$

$$Kc = \frac{[C]^c\ [D]^d}{[A]^a\ [B]^b}$$

Where:
[A] = concentration ($mol.dm^{-3}$)
a = no. of moles from equation

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COMMON IONS

POSITIVE

GROUP 1 = +
GROUP 2 = 2+
H^+
Ag^+
Zn^{2+}
Pb^{2+}
Al^{3+}
(Transition metals are variable)
e.g. Fe^{2+}, Fe^{3+}

NEGATIVE

GROUP 7 = -
GROUP 6 = 2-
GROUP 5 = 3-

MOLECULAR IONS

NH_4^+ ammonium	OH^- hydroxide	NO_3^- nitrate	CN^- cyanide
H_3O^+ hydronium	CO_3^{2-} carbonate	SO_4^{2-} sulfate	PO_4^{3-} phosphate

ACIDS & BASES

ACIDS

HCl	hydrochloric acid
HNO_3	nitric acid
H_2SO_4	sulphuric acid
H_3PO_4	phosphoric acid
CH_3COOH	ethanoic acid

BASES

$NaOH$	sodium hydroxide
KOH	potassium hydroxide
$Ca(OH)_2$	calcium hydroxide
CuO	copper (II) oxide

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COMMON OXIDATION STATES

POSITIVE

GROUP 1 = +I
GROUP 2 = +II
H = +I
Ag = +I
Zn = +II
Pb = +II or +IV
Al = + III
(Transition metals are variable)
Fe = +II or +III
Cu = +II (sometimes +I)
C = +II or +IV

NEGATIVE

F = -I
O = -II
Cl = -I
Br = -I
I = -I
N = -III
S = -II
P = -III

Most common oxidation states, but may be positive when covalently bonded to more highly electronegative elements.
i.e. F or O

GROUP 1 SALTS: ALL SOLUBLE

NITRATE SALTS = ALL SOLUBLE

GROUP 2 SALTS:
HYDROXIDES INCREASE IN SOLUBILITY DOWN THE GROUP
SULFATES DECREASE IN SOLUBILITY DOWN THE GROUP
CARBONATES ARE NOT SOLUBLE

Ag SALTS: ALL INSOLUBLE EXCEPT $AgNO_3$

Pb SALTS ALL INSOLUBLE EXCEPT $Pb(NO_3)_2$

GROUP 7 SALTS: ALL SOLUBLE EXCEPT AgX and PbX_2

CO_3 SALTS: ALL INSOLUBLE EXEPT GROUP 1

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YEAR 1 CHEMISTRY PRACTICALS (AQA)

No.	Practical	Detail	Done?
1a	Make a standard solution	Prepare a 250cm^3 sample of a solution (e.g. NaOH) to a known concentration using the standard method.	
1b	Perform a simple acid-base titration	Titrate an acid of known concentration against an alkali. Deduce the concentration of the alkali using a the mean titre. e.g. HCl + NaOH	
2a	Measure the enthalpy change of combustion of a fuel (ΔHc)	Use a calorimeter to experimentally determine the energy released by a fuel and the ΔH for the reaction.	
2b	Measure the enthalpy change of neutralisation (ΔHn)	Use a calorimeter to experimentally determine the energy released by a neutralisation reaction and the ΔH for the reaction.	
3	Measure temperature affects the rate of a reaction	Use the initial rate method to determine the effect of increasing temperature ion the rate of a reaction. e.g. HCl + Sodium Thiosufate	
4	Testing for ions in solution	Use chemical tests to identify Group 2, Group 7, OH^-, CO_3^{2-} and SO_4^{2-} ions in solution	
5	Distillation of an organic product	Produce a liquid organic compound and use distillation to separate and purify. e.g. Aldehyde or Ester	
6	Testing for organic functional groups	Use chemical tests to identify a carboxylic acid, an alcohol and an aldehyde.	

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THERMODYNAMICS

$$\Delta H^{\ominus}_{\text{solution}} = \underset{\text{(ENDOTHERMIC)}}{\Delta H^{\ominus}_{\text{latt diss}}} + \underset{\text{(EXOTHERMIC)}}{\sum \Delta H^{\ominus}_{\text{hydration}}}$$

$$\underset{(\text{J.K}^{-1}.\text{mol}^{-1})}{\Delta S_{\text{system}}} = \sum \Delta S_{\text{products}} - \sum \Delta S_{\text{reactants}}$$

$$\Delta G = \Delta H - T\,\Delta S$$

(kJ.mol^{-1})

Always divide ΔS by 1000 to match your units!

$$T = \frac{\Delta H}{\Delta S}$$

When calculating the temperature at which a reaction becomes / ceases being feasible

$$\Delta G = -\Delta S\,T + \Delta H$$

For the ΔG Vs T Graph!
Equivalent to:
$y = mx + c$

RATE EQUATIONS & ARRHENIUS

$$\underset{(\text{mol.dm}^{-3}.\text{s}^{-1})}{\text{Rate}} = k\,[A]^{\text{order}}\,[B]^{\text{order}}$$

$$K = \frac{\text{Rate}}{[A]^{\text{order}}\,[B]^{\text{order}}}$$

$$K = A\,e^{\left(\frac{-Ea}{RT}\right)} \quad \text{OR} \quad \ln K = \left(\frac{-Ea}{RT}\right) + \ln A \quad \text{OR} \quad \ln K = \frac{-Ea}{R} \times \frac{1}{T} + \ln A$$

Standard Arrhenius

Natural Log of Arrhenius

For the lnK Vs 1/T Graph!

Equivalent to:
$y = mx + c$

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ELECTROCHEMISTRY

$$\text{EMF}_{(V)} = \text{Most Positive } E^{\circ} - \text{Most Negative } E^{\circ}$$

OR

$$\text{EMF} = E^{\circ} \text{ Cell being Reduced} - E^{\circ} \text{ Cell being Oxidised}$$

ACIDS & BASES

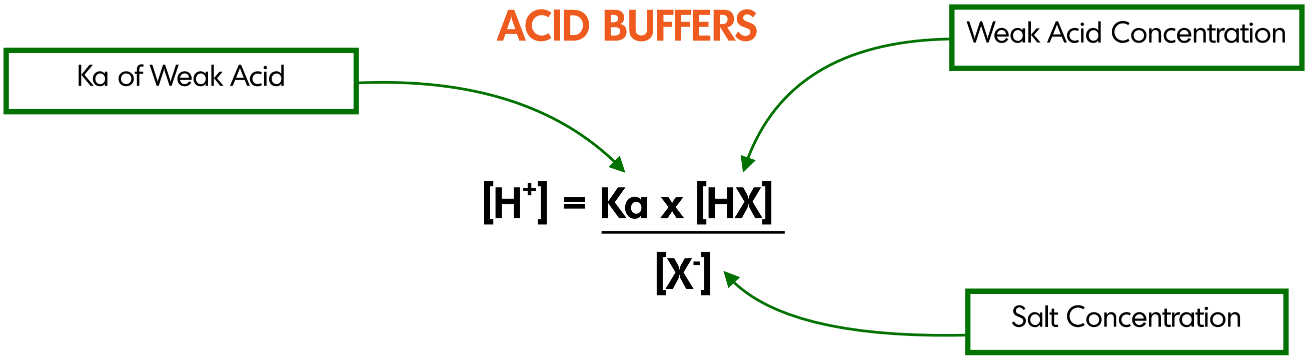
$$pH = -\log [H^{+}] \quad \text{and} \quad [H^{+}] = 10^{-pH}$$

$$Ka = \frac{[H^{+}]\,[X^{-}]}{[HX]}$$

$$Kw = [H^{+}]\,[OH^{-}]$$

$Kw = 1.00 \times 10^{-14}\ mol^{2}.dm^{-6}$
at 298K

ACID BUFFERS



Ka of Weak Acid

Weak Acid Concentration

$$[H^{+}] = \frac{Ka \times [HX]}{[X^{-}]}$$

Salt Concentration

Then $pH = -\log [H^{+}]$

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YEAR 2 TRANSITION METALS (AQA)

COMMON OXIDATION STATES

Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
+III	+IV	+II, +III +IV, +V	+II, +III +VI	+II, +III +IV, +VI +VII	+II, +III	+II, +III	+II	+I, +II	+II

COMMON COMPLEX COLOURS

	$+H_2O$	$+ OH^-_{(aq)}$ or limited $NH_{3(aq)}$	+ Excess $NH_{3(aq)}$	+ conc. $HCl_{(aq)}$	$+ CO_3^{2-}{}_{(aq)}$
Iron II	$[Fe(H_2O)_6]^{2+}{}_{(aq)}$	$[Fe(H_2O)_4(OH)_2]_{(s)}$	NVR	NVR	$FeCO_{3\,(s)}$
	GREEN SOLUTION	GREEN PRECIPITATE (may oxidise to brown)			GREEN PRECIPITATE
Iron III	$[Fe(H_2O)_6]^{3+}{}_{(aq)}$	$[Fe(H_2O)_3(OH)_3]_{(s)}$	NVR	$FeCl_4^-$	$[Fe(H_2O)_3(OH)_3]_{(s)}$
	YELLOW SOLUTION	BROWN PRECIPITATE		YELLOW SOLUTION	BROWN PRECIPITATE & $CO_2(g)$
***Cobalt**	$[Co(H_2O)_6]^{2+}{}_{(aq)}$	$[Co(H_2O)_4(OH)_2]_{(s)}$	$[Co(NH_3)_6]^{2+}{}_{(aq)}$	$CoCl_4^-$	$CoCO_{3\,(s)}$
	PINK SOLUTION	GREEN PRECIPITATE	BROWN SOLUTION	BLUE SOLUTION	PURPLE PRECIPITATE
Copper	$[Cu(H_2O)_6]^{2+}{}_{(aq)}$	$[Cu(H_2O)_4(OH)_2]_{(s)}$	$[Cu(H_2O)_2(NH_3)_4]^{2+}{}_{(aq)}$	$CuCl_4^{2-}$	$CuCO_{3\,(s)}$
	LIGHT BLUE SOLUTION	BLUE PRECIPITATE	ROYAL BLUE SOLUTION	YELLOW / GREEN SOLUTION	BLUE / GREEN PRECIPITATE
Aluminum	$[Al(H_2O)_6]^{3+}{}_{(aq)}$	$[Al(H_2O)_3(OH)_3]_{(s)}$	NVR	NVR	$[Al(H_2O)_3(OH)_3]_{(s)}$
	COLOURLESS SOLUTION	WHITE PRECIPITATE			WHITE PRECIPITATE & $CO_2(g)$

* Not strictly on the spec but has been known to appear in exam

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VANADIUM

H^+ / Zn can reduce vanadium in $VO_3^-{}_{(aq)}$

Species	VO_3^-	VO_2^+	VO^{2+}	V^{3+}	V^{2+}
Oxidation State	+V	+V	+IV	+III	+II
Colour of solution	YELLOW	YELLOW	BLUE	GREEN	PURPLE

WHY COMPLEXES ARE COLOURED

$$\triangle E = h.f$$

$\triangle$E = Difference in energy between d orbitals

h = Planck's Constant (6.63×10^{-34} J.S^{-1}) **f** = Frequency (Hz)

$$c = \lambda.f$$

c = Speed of light (3.00×10^{8} m.s^{-1})

λ = Wavelength (nm) **f** = Frequency (Hz)

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No.	Practical	Detail	Done?
7	Measure rate by initial rate method	The "Iodine Clock" Reaction Iodide (V) + H_2O_2	
8	Set up an electrochemical cell	Set up an electrochemical cell and measure EMF	
9	pH Curve	Investigate how pH changes when a weak acid reacts with a strong base	
10a	Preparation of an organic solid	Produce a pure sample of an organic solid (e.g. Aspirin) and test its purity	
10b	Preparation of an organic liquid	Produce a pure sample of an organic liquid (e.g. ethyl ethanoate) and test its purity	
11	Testing for ions in solution	Use chemical tests to identify transition metal ions in solution	
12	Thin Layer Chromatography	Use TLC to analyse the composition of mixtures of compounds	

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