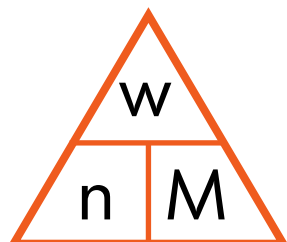


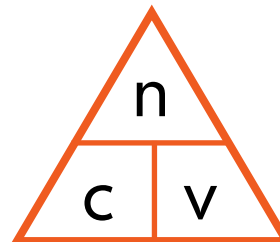


MOLE CALCULATIONS

number of moles = mass / molar mass
(g) (g.mol⁻¹)



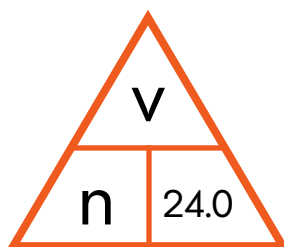
number of moles = concentration x volume
(mol.dm⁻³) (dm³)



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

MOLAR GAS CONSTANT

1 mole of ANY gas occupies 24.0 dm³ at room temperature & pressure



IDEAL GAS EQUATION

P = Pressure (pa) **V** = volume (m³) **n** = no. of moles
R = Gas Constant (8.314 J.K⁻¹.mol⁻¹) **T** = Temperature (K)

$$PV = nRT$$

$$P = \frac{nRT}{V}$$

$$V = \frac{nRT}{P}$$

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

$$T = \frac{PV}{nR}$$

For changes in conditions:

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



MASS SPECTROSCOPY

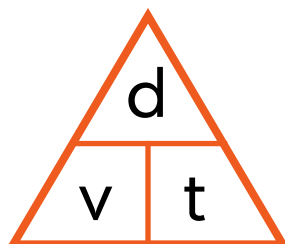
K.E. = Kinetic Energy (J) m = mass of ONE ion (Kg) v = velocity (m.s⁻¹)

$$\text{K.E.} = \frac{1}{2} m.v^2$$

$$v = \frac{\sqrt{2\text{K.E.}}}{m}$$

$$m = \frac{\sqrt{2\text{K.E.}}}{v}$$

d = distance or length of flight tube (m) t = time (s) v = velocity (m.s⁻¹)



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

OTHER EQUATIONS

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

$$\text{Empirical formula} = \frac{M_1}{M_{r1}} : \frac{M_2}{M_{r2}} : \frac{M_3}{M_{r3}}$$

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!



COMMON IONS

POSITIVE

GROUP 1 = +

GROUP 2 = 2+

H⁺

Ag⁺

Zn²⁺

Pb²⁺

Al³⁺

(Transition metals are variable)

e.g. Fe²⁺, Fe³⁺

NEGATIVE

GROUP 7 = -

GROUP 6 = 2-

GROUP 5 = 3-

MOLECULAR IONS

NH₄⁺
ammonium

H₃O⁺
hydronium

OH⁻
hydroxide

CO₃²⁻
carbonate

NO₃⁻
nitrate

SO₄²⁻
sulfate

CN⁻
cyanide

PO₄³⁻
phosphate

ACIDS & BASES

ACIDS

HCl hydrochloric acid

HNO₃ nitric acid

H₂SO₄ sulphuric acid

H₃PO₄ phosphoric acid

CH₃COOH ethanoic acid

BASES

NaOH sodium hydroxide

KOH potassium hydroxide

Ca(OH)₂ calcium hydroxide

CuO copper (II) oxide



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 $\text{Pb}(\text{NO}_3)_2$

GROUP 7 SALTS: ALL SOLUBLE EXCEPT AgX and PbX_2

CO_3 SALTS: ALL INSOLUBLE EXCEPT GROUP 1



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 (ΔH_c)	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 (ΔH_n)	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.	



THERMODYNAMICS

$$\Delta H^{\ominus}_{\text{solution}} = \Delta H^{\ominus}_{\text{latt diss}} + \sum \Delta H^{\ominus}_{\text{hydration}}$$

(ENDOTHERMIC) (EXOTHERMIC)

$$\Delta S_{\text{system}} = \sum \Delta S_{\text{products}} - \sum \Delta S_{\text{reactants}}$$

(J.K⁻¹.mol⁻¹)

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

(kJ.mol⁻¹)

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

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

Always divide ΔS by 1000
to match your units!

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

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

RATE EQUATIONS & ARRHENIUS

$$\text{Rate} = k [\text{A}]^{\text{order}} [\text{B}]^{\text{order}}$$

(mol.dm⁻³.s⁻¹)

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

$$K = A e^{\left(\frac{-E_a}{RT}\right)}$$

OR

$$\ln K = \left(\frac{-E_a}{RT}\right) + \ln A$$

OR

$$\ln K = \frac{-E_a}{R} \times \frac{1}{T} + \ln A$$

Standard Arrhenius

Natural Log of
Arrhenius

For the $\ln K$ Vs $1/T$ Graph!

Equivalent to:
 $y = mx + c$



ELECTROCHEMISTRY

$$\text{EMF} = \text{Most Positive } E^\circ - \text{Most Negative } E^\circ$$

(M)

OR

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

ACIDS & BASES

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

$$K_a = \frac{[\text{H}^+] [\text{X}^-]}{[\text{HX}]}$$

$$K_w = [\text{H}^+] [\text{OH}^-]$$

$$K_w = 1.00 \times 10^{-14} \text{ mol}^2 \cdot \text{dm}^{-6} \text{ at } 298\text{K}$$

ACID BUFFERS

K_a of Weak Acid

Weak Acid Concentration

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

Salt Concentration

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



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 ₂ O	+ OH ⁻ _(aq) or limited NH _{3(aq)}	+ Excess NH _{3(aq)}	+ conc. HCl _(aq)	+ CO ₃ ²⁻ _(aq)
Iron II	[Fe(H ₂ O) ₆] ²⁺ _(aq) GREEN SOLUTION	[Fe(H ₂ O) ₄ (OH) ₂] _(s) GREEN PRECIPITATE (may oxidise to brown)	NVR	NVR	FeCO ₃ (s) GREEN PRECIPITATE
	[Fe(H ₂ O) ₆] ³⁺ _(aq) YELLOW SOLUTION	[Fe(H ₂ O) ₃ (OH) ₃] _(s) BROWN PRECIPITATE	NVR	FeCl ₄ ⁻ YELLOW SOLUTION	[Fe(H ₂ O) ₃ (OH) ₃] _(s) BROWN PRECIPITATE & CO ₂ (g)
*Cobalt	[Co(H ₂ O) ₆] ²⁺ _(aq) PINK SOLUTION	[Co(H ₂ O) ₄ (OH) ₂] _(s) GREEN PRECIPITATE	[Co(NH ₃) ₆] ²⁺ _(aq) BROWN SOLUTION	CoCl ₄ ⁻ BLUE SOLUTION	CoCO ₃ (s) PURPLE PRECIPITATE
	[Cu(H ₂ O) ₆] ²⁺ _(aq) LIGHT BLUE SOLUTION	[Cu(H ₂ O) ₄ (OH) ₂] _(s) BLUE PRECIPITATE	[Cu(H ₂ O) ₂ (NH ₃) ₄] ²⁺ _(aq) ROYAL BLUE SOLUTION	CuCl ₄ ²⁻ YELLOW / GREEN SOLUTION	CuCO ₃ (s) BLUE / GREEN PRECIPITATE
Aluminum	[Al(H ₂ O) ₆] ³⁺ _(aq) COLOURLESS SOLUTION	[Al(H ₂ O) ₃ (OH) ₃] _(s) WHITE PRECIPITATE	NVR	NVR	[Al(H ₂ O) ₃ (OH) ₃] _(s) WHITE PRECIPITATE & CO ₂ (g)

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



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

$$\Delta E = h.f$$

ΔE = Difference in energy between d orbitals

h = Planck's Constant ($6.63 \times 10^{-34} \text{ J}\cdot\text{s}^{-1}$) f = Frequency (Hz)

$$c = \lambda.f$$

c = Speed of light ($3.00 \times 10^8 \text{ m}\cdot\text{s}^{-1}$)

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



No.	Practical	Detail	Done?
7	Measure rate by initial rate method	The "Iodine Clock" Reaction Iodide (I ⁻) + H ₂ O 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	