Must-know **A level chemistry** equations

Number of moles

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

n - number of moles [mol]
m - mass [g]
M - molar mass [g mol⁻¹]

Number of moles for gases

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

n - number of moles [mol]
V - volume [dm³]
V_m - molar volume

Note:

V_m = 22.4 OR 22.7 dm³ mol⁻¹depending on how your curriculum defines STP (standard temperature and pressure)

V_m = 24 dm³ mol⁻¹ at RTP (room temperature and pressure)

Molar concentration

$$c = \frac{n}{V_{solution}}$$

c - concentration [mol dm⁻³]
n - number of moles [mol]
$V_{solution}$ - volume of solution [dm³]
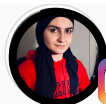
Ideal gas law

$$pV = nRT$$

p - pressure [Pa]
V - volume [m³]
n - number of moles [mol]
R - gas constant [8.314 J mol⁻¹ K⁻¹]
T - temperature [K]

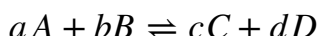## Heat released or absorbed

$$Q = mc\Delta T$$

Q - amount of heat [J]
m - mass [g]
c - specific heat capacity [for water: 4.18 J g⁻¹ K⁻¹]
ΔT - temperature change [K or °C]

The equilibrium constant K_c

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

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

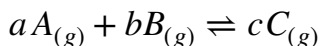[] - concentration at equilibrium [mol dm⁻³]

Note:

Acid dissociation constant K_a and base dissociation constant K_b can be written in the same way as K_c.

For example, dissociation constant of a weak acid:

$$HA \rightleftharpoons H^+ + A^-$$

$$K_a = \frac{[H^+] \cdot [A^-]}{[HA]}$$

The equilibrium constant K_p (for gases)

$$aA_{(g)} + bB_{(g)} \rightleftharpoons cC_{(g)}$$

$$K_p = \frac{p(C)^c}{p(A)^a \cdot p(B)^b}$$

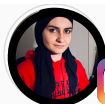$p()$ - partial pressure [Pa, or kPa, or atm]

Partial pressure

$$p(A) = \chi(A) \cdot p_{total}$$

$p(A)$ - partial pressure of gas A
$\chi(A)$ - mole fraction of gas A
p_{total} - total pressure

Mole fraction

$$\chi(A) = \frac{n(A)}{n_{total}}$$

$n(A)$ - number of moles of gas A
n_{total} - total number of moles of gases

pH

$$pH = -log_{10}[H^+]$$

$[H^+]$ - concentration of H^+ ions [mol dm⁻³]

The amount of charge

$$Q = It$$

Q - quantity of charge [C]
I - current [A]
t - time [s]

Nernst equation for a metal dipping into a solution of its salt

$$E = E^0 + \frac{2.3RT}{zF}log_{10}[ion]$$

E - electrode potential [V]
E^0 - standard electrode potential [V]
R - gas constant [8.314 J mol⁻¹ K⁻¹]
T - temperature [K]
z - number of electrons involved in the reaction
F - Faraday constant [96500 C mol⁻¹]
$[ion]$ - ion concentration [mol dm⁻³]

Nernst equation for a system with two ions

$$E = E^0 + \frac{2.3RT}{zF}log_{10}\frac{[oxidized]}{[reduced]}$$
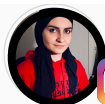
E - electrode potential [V]
E^0 - standard electrode potential [V]
R - gas constant [8.314 J mol⁻¹ K⁻¹]
T - temperature [K]
z - number of electrons transferred
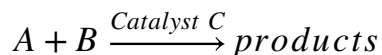F - Faraday constant [96500 C mol⁻¹]

[*oxidized*] - concentration of oxidized form [mol dm⁻³]

[*reduced*] - concentration of reduced form [mol dm⁻³]

Rate constant

$$A + B \xrightarrow{Catalyst\ C} products$$

$$r = k[A]^x[B]^y[C]^z$$

r - reaction rate

k - rate constant

[] - concentration [mol dm⁻³]

x, y, z - orders of reaction with respect to A, B, C

Note:

x, y, z are defined experimentally! These are not the coefficients in the equation.

Entropy

$$\Delta S_{total} = \Delta S_{surr} + \Delta S_{reaction}$$

ΔS_{total} - total entropy change

ΔS_{surr} - entropy change of surroundings

$\Delta S_{reaction}$ - entropy change of reaction

$$\Delta S_{surr} = -\frac{\Delta H}{T}$$

ΔS_{surr} - entropy change of surroundings [J mol⁻¹ K⁻¹]

ΔH - enthalpy change of reaction [J mol⁻¹]

T - temperature [K]

Gibbs free energy

$$\Delta G = \Delta H - T\Delta S_{reaction}$$

ΔG - Gibbs free energy [J mol⁻¹]

$\Delta S_{reaction}$ - entropy change of reaction [J mol⁻¹ K⁻¹]

ΔH - enthalpy change of reaction [J mol⁻¹]

T - temperature [K]