

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⁻¹ depending on how your curriculum defines STP (standard temperature and pressure)

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

Molar concentration

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

c - concentration [mol dm⁻³]

n - number of moles [mol]

V_{solution} - volume of solution [dm³]

Ideal gas law

$$pV = nRT$$

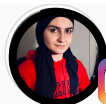
p - pressure [Pa]

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

$$Q = mc\Delta T$$

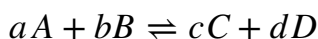
Q - amount of heat [J]

m - mass [g]

c - specific heat capacity [for water: $4.18 \text{ J g}^{-1} \text{ K}^{-1}$]

ΔT - temperature change [K or $^{\circ}\text{C}$]

The equilibrium constant K_c



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

[] - concentration at equilibrium [mol dm^{-3}]

Note:

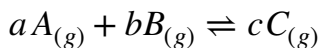
Acid dissociation constant K_a and base dissociation constant K_b can be written in the same way as K_c .

For example, dissociation constant of a weak acid:



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

The equilibrium constant K_p (for gases)



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

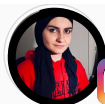
$p()$ - partial pressure [Pa, or kPa, or atm]

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 [$8.314 \text{ J mol}^{-1} \text{ K}^{-1}$]

T - temperature [K]

z - number of electrons involved in the reaction

F - Faraday constant [96500 C mol^{-1}]

$[ion]$ - ion concentration [mol dm^{-3}]

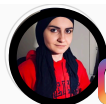
Nernst equation for a system with two ions

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

E - electrode potential [V]

E^0 - standard electrode potential [V]

R - gas constant [$8.314 \text{ J mol}^{-1} \text{ K}^{-1}$]



T - temperature [K]

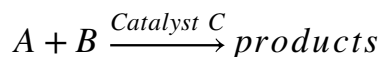
z - number of electrons transferred

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



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

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

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

ΔS_{total} - total entropy change

ΔS_{surr} - entropy change of surroundings

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

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

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

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

ΔG - Gibbs free energy [J mol⁻¹]

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

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

T - temperature [K]

