# 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 [dm$^3$]
$V_m$ - molar volume

Note:

$V_m$ = 22.4 OR 22.7 dm$^3$ mol$^{-1}$depending on how your curriculum defines STP (standard temperature and pressure)

$V_m$ = 24 dm$^3$ 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 [dm$^3$]

## Ideal gas law

$$pV = nRT$$

$p$ - pressure [Pa]
$V$ - volume [m$^3$]
$n$ - 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]
$c$ - specific heat capacity [for water: 4.18 J $g^{-1}$ $K^{-1}$]
$\Delta T$ - 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

## Entropy

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

$\Delta S_{total}$ - total entropy change
$\Delta S_{surr}$ - entropy change of surroundings
$\Delta S_{reaction}$ - entropy change of reaction

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

$\Delta S_{surr}$ - entropy change of surroundings [J mol$^{-1}$ K$^{-1}$]
$\Delta H$ - enthalpy change of reaction [J mol$^{-1}$]
$T$ - temperature [K]

## Gibbs free energy

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

$\Delta G$ - Gibbs free energy [J mol$^{-1}$]
$\Delta S_{reaction}$ - entropy change of reaction[J mol$^{-1}$ K$^{-1}$]
$\Delta H$ - enthalpy change of reaction [J mol$^{-1}$]
$T$ - temperature [K]