

Equilibrium: one-page cheat sheet

NEET Chemistry · Unit 6 · Physical Chemistry

1 • Equilibrium constant



Reverse $K_{\text{rev}} = 1/K$

*** n** $K' = K^n$

Add equations $K_{\text{net}} = K_1 \times K_2$

$$K_p = K_c (RT)^{\Delta n_g}$$

$\Delta n_g = 0$ $K_p = K_c$ ($H_2 + I_2, N_2 + O_2$)

$\Delta n_g > 0$ $K_p > K_c$ (PCl_5 dissociation)

$\Delta n_g < 0$ $K_p < K_c$ ($N_2 + 3H_2$)

2 • Reaction quotient

Compare	Result
$Q_c < K_c$	forward reaction
$Q_c = K_c$	at equilibrium
$Q_c > K_c$	backward reaction

$$\Delta G = \Delta G^\circ + RT \ln Q, \quad \Delta G^\circ = -2.303 RT \log K$$

3 • Le Chatelier's principle

Change	Shift
add reactant / remove product	forward
add product	backward
increase pressure	fewer gas moles
decrease pressure	more gas moles
increase T	endothermic direction
decrease T	exothermic direction
catalyst	no shift; faster approach
inert gas, constant V	no shift
inert gas, constant P	more gas moles

Trap • Only **temperature** changes the value of K . Concentration, pressure and catalyst do not.

4 • Acids, bases & water

Arrhenius acid gives H^+ , base gives OH^-

Brønsted–Lowry acid donates H^+ , base accepts; conjugate pair differs by one H^+

Lewis acid accepts an electron pair ($BF_3, AlCl_3$); base donates one (NH_3)

$$K_w = [H^+][OH^-] = 10^{-14}, \quad pH + pOH = 14, \quad K_a K_b = K_w$$

K_w rises with T , so neutral pH falls below 7 on heating.

5 • Weak electrolytes

$$K_a = \frac{C\alpha^2}{1-\alpha} \Rightarrow \alpha = \sqrt{\frac{K_a}{C}}, \quad [H^+] = \sqrt{K_a C}$$

$$pH = \frac{1}{2} (pK_a - \log C)$$

Common ion CH_3COONa added to CH_3COOH lowers $[H^+]$, raises pH

6 • Hydrolysis & buffers

Salt of	pH at 298 K
weak acid + strong base	$7 + \frac{1}{2}(pK_a + \log C)$
strong acid + weak base	$7 - \frac{1}{2}(pK_b + \log C)$
weak acid + weak base	$7 + \frac{1}{2}(pK_a - pK_b)$

Acidic buffer $pH = pK_a + \log \frac{[\text{salt}]}{[\text{acid}]}$

Basic buffer $pOH = pK_b + \log \frac{[\text{salt}]}{[\text{base}]}$

Trap • Weak acid + weak base salt: pH does **not** depend on concentration.

7 • Solubility product

Type	K_{sp}	Examples
AB	s^2	AgCl, $BaSO_4$
AB_2 or A_2B	$4s^3$	$Mg(OH)_2, Ag_2CrO_4$
AB_3	$27s^4$	$Fe(OH)_3, Al(OH)_3$
A_2B_3	$108s^5$	As_2S_3
A_3B_4	$6912s^7$	$Zr_3(PO_4)_4$

Precipitates when ionic product $> K_{sp}$

Group II H_2S in dilute HCl: common H^+ keeps $[S^{2-}]$ low; only CuS, PbS, Bi_2S_3 settle

Group III NH_4OH with NH_4Cl : common NH_4^+ limits $[OH^-]$; $Fe(OH)_3, Al(OH)_3$ settle

8 • Quick pH rules

Strong acid $pH = -\log C$ (monoprotic)

Strong base $pOH = -\log C$, then $pH = 14 - pOH$

Acid strength larger K_a means smaller pK_a means stronger acid

Solubility a common ion **lowers** the solubility of a sparingly soluble salt

Seen in NEET 2026

Buffer with equal $[X^-]$ and $[HX]$; $K_b(X^-) = 10^{-10}$. Find pH.

$$K_a = \frac{10^{-14}}{10^{-10}} = 10^{-4} \Rightarrow pH = pK_a + \log 1 = 4$$

Seen in NEET 2022

50 mL each of 0.10 M sodium acetate and 0.01 M acetic acid; $pK_a = 4.57$.

$$pH = 4.57 + \log \frac{0.10}{0.01} = 5.57$$

Six traps to remember

- Only temperature changes K .
- $K_p = K_c(RT)^{\Delta n_g}$; count **gas** moles only.
- $Q > K$ means the reaction runs **backward**.
- A catalyst never shifts the equilibrium; it only speeds up reaching it.
- Inert gas at constant volume does nothing; at constant pressure it favours more gas moles.
- Henderson equation uses $\log(\text{salt}/\text{acid})$; equal amounts give $pH = pK_a$.