06

JEE MAIN · PAPER 1 · CHEMISTRY

Equilibrium

Formula & concept sheet

Why reactions stop short of completion and how to predict where: dynamic equilibrium, K_c and K_p, Le Chatelier's principle, then ionic equilibrium (acids and bases, pH, hydrolysis, buffers, titration curves) and solubility product. Each idea is explained first, then the formulas, shortcuts and traps.

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READ ALONGSIDE NCERT XI · Equilibrium

HOW TO USE THIS SHEET

1 **Skim**
Read the Key ideas, then the formula cards.

2 **Memorise**
Learn the formulas and check every Trap box.

3 **Test**
Do the worked examples, then this chapter's PYQs.

4 **Revise**
Use the last-minute checklist before the exam.

Formula memorise · **Key idea** why it works · **Trap** marks lost here · **PYQ** how it is asked

Practise this chapter with real past-paper questions on Prep Vault. Written for revision, not as a first read.

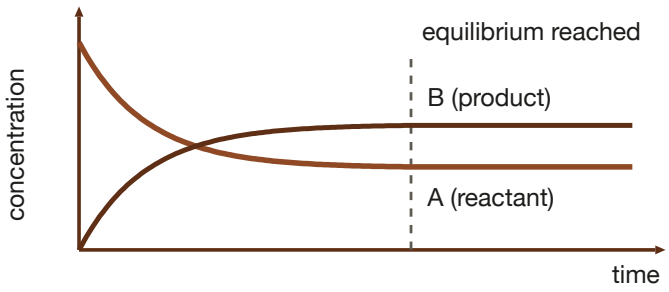
CHAPTER 06

Equilibrium

NCERT XI · Equilibrium

Covers: dynamic equilibrium and physical equilibria · law of chemical equilibrium, K_c and K_p · ΔG, ΔG° and K · Le Chatelier's principle · acids and bases (Arrhenius, Brønsted–Lowry, Lewis) · ionisation of water, pH · weak acids and bases, multistage ionisation · common ion effect · salt hydrolysis · buffers · solubility product.

1 Dynamic equilibrium and the equilibrium constant



forward and reverse rates become equal, so the concentrations stop changing

- Many reactions do not go to completion. When forward and reverse rates become equal, the concentrations stop changing: this is **dynamic equilibrium**. Nothing visible happens, but both reactions continue.
- Equilibrium can be approached from either side and is reached only in a **closed** system. At a given temperature the same ratio of concentrations (the constant K) is reached from reactants alone, products alone, or any mixture.
- A **catalyst** speeds up both directions equally: equilibrium is reached sooner, but its position and K do not change.

PHYSICAL EQUILIBRIA

EQUILIBRIUM	EXAMPLE AND FACT
solid $\rightleftharpoons$ liquid	ice $\rightleftharpoons$ water at 273 K and 1 atm; higher pressure lowers the melting point of ice (water is denser than ice)
liquid $\rightleftharpoons$ vapour	water in a closed vessel: vapour pressure is constant at a given T, rises with T
solid $\rightleftharpoons$ vapour	sublimation of iodine or camphor in a closed vessel
gas $\rightleftharpoons$ dissolved gas	Henry's law: $p = K_H x$ (see Solutions); solubility falls as temperature rises

LAW OF CHEMICAL EQUILIBRIUM For $a\,\mathrm{A} + b\,\mathrm{B} \rightleftharpoons c\,\mathrm{C} + d\,\mathrm{D}$ at equilibrium and constant temperature, $K_c = [C]^c[D]^d / ([A]^a[B]^b)$, using molar concentrations; K_p is the same expression with partial pressures. Pure solids and pure liquids (and the solvent, water) are left out: their activity is 1.

K_p AND K_c

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

Δn_g = gaseous products − gaseous reactants; $R = 0.0821$ L atm K^{-1} mol^{-1} with p in atm (0.083 with bar)

MOLE FRACTIONS

$$K_p = K_x\,P^{\Delta n_g}$$

K_x from mole fractions, P the total pressure; $p_i = x_i P$

DIRECTION: REACTION QUOTIENT

$Q < K$ forward, $Q > K$ backward

Q has the form of K with the present concentrations; $Q = K$ at equilibrium

MANIPULATING K

CHANGE TO THE EQUATION	NEW CONSTANT
reverse the reaction	$1/K$
multiply the coefficients by n	K^n
add two equations	$K_1 \times K_2$
subtract (reverse and add)	K_1/K_2
change of temperature	K changes (the only thing that changes K)

KEY IDEA

$K \gg 1$: mostly products at equilibrium; $K \ll 1$: mostly reactants; K says how far, not how fast. K depends only on the equation as written and on T, never on concentrations, pressure, volume or catalyst.

TRAP

$K_p = K_c$ only if $\Delta n_g = 0$ ($\mathrm{H_2 + I_2 \rightleftharpoons 2\,HI}$). For $\mathrm{CaCO_3\,(s) \rightleftharpoons CaO\,(s) + CO_2\,(g)}$, $K_p = p_{\mathrm{CO_2}}$ and $K_c = [\mathrm{CO_2}]$: the solids are left out. Keep the unit of p consistent with R.

2 Degree of dissociation and equilibrium calculations

THE ICE TABLE

1. Write the balanced equation and the initial moles (or concentrations) of every species.
2. Let the change be $-x$ (or $-\alpha\times$ initial) for the reactant; products change by their coefficients times x.
3. Equilibrium = initial + change. For gases, find the total moles n_t and use $p_i = \dfrac{n_i}{n_t}P$ (or divide moles by V for K_c).
4. Substitute into K_p or K_c, solve, and check that the result is physical (moles > 0).

STANDARD DISSOCIATION RESULTS (START WITH 1 MOL, DEGREE OF DISSOCIATION α, TOTAL PRESSURE P)

EQUILIBRIUM	TOTAL MOLES	EXPRESSION
$\mathrm{PCl_5 \rightleftharpoons PCl_3 + Cl_2}$	$1+\alpha$	$K_p = \dfrac{\alpha^2 P}{1-\alpha^2}$
$\mathrm{N_2O_4 \rightleftharpoons 2\,NO_2}$	$1+\alpha$	$K_p = \dfrac{4\alpha^2 P}{1-\alpha^2}$
$\mathrm{2\,NH_3 \rightleftharpoons N_2 + 3\,H_2}$	$1+\alpha$	$K_p = \dfrac{27\alpha^4 P^2}{16(1-\alpha^2)^2}$
$\mathrm{A \rightleftharpoons B + \tfrac{1}{2}\,C}$	$1+\alpha/2$	$K_p = \dfrac{\alpha\,(\alpha/2)^{1/2}P^{1/2}}{(1-\alpha)(1+\alpha/2)^{1/2}}$ (derive; small α: $\dfrac{\alpha^{3/2}P^{1/2}}{\sqrt{2}}$)

DEGREE OF DISSOCIATION FROM VAPOUR DENSITY

$$\alpha = \frac{D-d}{(n-1)\,d}$$

D = initial (undissociated) vapour density, d = vapour density at equilibrium, n = moles of gas from 1 mol of reactant; $d \propto 1/n_t$

LINK TO GIBBS ENERGY

$$\Delta G^\circ = -2.303\,RT\log K$$

$\Delta G = \Delta G^\circ + 2.303\,RT\log Q$; $K = 10^{-\Delta G^\circ/2.303RT}$ (see Thermodynamics)

TEMPERATURE DEPENDENCE (VAN'T HOFF)

$$\log\frac{K_2}{K_1} = \frac{\Delta H^\circ}{2.303R}\left(\frac{1}{T_1}-\frac{1}{T_2}\right)$$

ΔH° in J mol^{-1}; endothermic: K rises with T; exothermic: K falls

TRAP

Raising the temperature speeds up **both** directions. K rises only if the forward reaction is endothermic. Do not use $\alpha \ll 1$ shortcuts when α is not small (above about 5%): solve the quadratic.

QUICK CHECK

PCl_5 is 50% dissociated at 2 atm total pressure. K_p?

Answer: $K_p = \dfrac{\alpha^2 P}{1-\alpha^2} = \dfrac{0.25 \times 2}{0.75} = 0.667$ atm.

QUICK CHECK

$K_c = 0.5$ for $N_2 + 3\,H_2 \rightleftharpoons 2\,NH_3$ at 500 K. K_p (atm)?

Answer: $\Delta n_g = -2$: $K_p = \dfrac{0.5}{(0.0821 \times 500)^2} = 3.0 \times 10^{-4}$.

3 Le Chatelier's principle

LE CHATELIER'S PRINCIPLE If a system at equilibrium is disturbed by a change of concentration, pressure, volume or temperature, it shifts in the direction that partly opposes the change. It tells the direction of the shift, not its size.

EFFECT OF A CHANGE ON $A + B \rightleftharpoons C + D$-TYPE EQUILIBRIA

CHANGE	SHIFT	DOES K CHANGE?
add reactant or remove product	forward	no
add product or remove reactant	backward	no
increase pressure (decrease volume)	towards fewer gas moles	no
no change in gas moles ($\Delta n_g = 0$)	no shift with pressure	no
inert gas, constant volume	none (partial pressures unchanged)	no
inert gas, constant pressure	towards more gas moles (volume grows, like a pressure drop)	no
raise T: forward endothermic	forward	K increases
raise T: forward exothermic	backward	K decreases
catalyst	no shift; equilibrium reached faster	no

INDUSTRIAL EXAMPLES

- Haber: $N_2 + 3\,H_2 \rightleftharpoons 2\,NH_3$, $\Delta H \approx -92$ kJ: high pressure favours NH_3; low T favours yield, but about 700 K and iron catalyst is the compromise for rate.
- Contact: $2\,SO_2 + O_2 \rightleftharpoons 2\,SO_3$, exothermic: high pressure, about 720 K, V_2O_5.
- $CaCO_3 \rightleftharpoons CaO + CO_2$: endothermic, gas formed; heated in an open kiln so CO_2 escapes and the reaction goes on.

WHY IT WORKS

The system re-balances Q against K. Adding a reactant makes $Q < K$, so the forward reaction runs until $Q = K$ again. A temperature rise is different: it changes K itself, in the direction that absorbs heat.

TRAP

Adding an inert gas at **constant volume** does nothing. A catalyst never shifts equilibrium or changes K. Increasing pressure by adding more of a reactant is a concentration change, not a volume change. Heating always raises both rates.

QUICK CHECK

Effect of raising T and of compressing on $N_2O_4 \rightleftharpoons 2\,NO_2$ ($\Delta H = +57$ kJ)?

Answer: Endothermic: higher T shifts forward (more brown NO_2). Compression shifts to fewer moles: backward (paler).

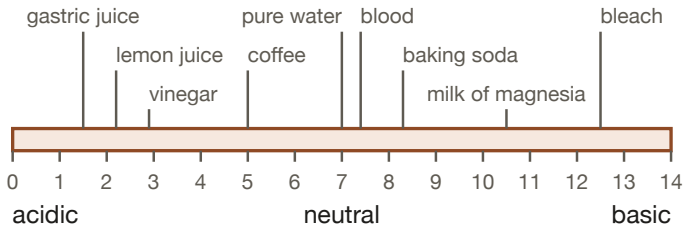
4 Acids, bases, pH and the ionisation of water

ACID-BASE DEFINITIONS

- **Arrhenius**: acid gives H^+, base gives OH^- in water.
- **Brønsted-Lowry**: acid is a proton donor, base a proton acceptor; an acid and its **conjugate base** differ by one H^+. A stronger acid has a weaker conjugate base.
- **Lewis**: acid is an electron-pair acceptor (BF_3, $AlCl_3$, Fe^{3+}, H^+, carbocations), base an electron-pair donor (NH_3, H_2O, OH^-, amines).

EXAMPLES AND AMPHOTERIC SPECIES

- $HCl + H_2O \longrightarrow H_3O^+ + Cl^-$: HCl acid, Cl^- its conjugate base.
- $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$: NH_3 base, NH_4^+ its conjugate acid.
- Amphoteric (both): H_2O, HCO_3^-, HSO_4^-, $H_2PO_4^-$.
- Conjugate pair constants: $K_aK_b = K_w$, $pK_a + pK_b = 14$.



$\mathrm{pH} = -\log[H^+]$; each unit is a factor of 10

IONIC PRODUCT OF WATER

$K_w = [H^+][OH^-] = 10^{-14}$ (at 298 K)

$pH + pOH = 14$; K_w rises with T, so hot pure water is neutral at pH below 7 (about 6.1 at 373 K, $K_w \approx 5.5 \times 10^{-13}$)

PH

$pH = -\log[H^+]$

pH < 7 acidic, $= 7$ neutral, > 7 basic (at 298 K); a change of one pH unit is a factor 10

STRENGTH OF ACIDS AND BASES

IDEA	RULE
strong acids	HCl, HBr, HI, HNO_3, H_2SO_4 (first step), $HClO_4$: fully ionised, $[H^+] = C$
strong bases	$NaOH$, KOH, $Ba(OH)_2$, $Ca(OH)_2$: $[OH^-] = C\times$ number of OH
very dilute strong acid	for $C < 10^{-6}$ M, include water: $[H^+] = C + [OH^-]$; 10^{-8} M HCl gives pH 6.98, not 8
acid strength and K_a	larger K_a (smaller pK_a) is stronger; oxo-acids: more oxygens on the central atom, stronger ($HClO_4 > HClO_3 > HClO_2 > HClO$)

WEAK ACID (OSTWALD)

$$K_a = \frac{C\alpha^2}{1-\alpha} \approx C\alpha^2$$

$\alpha = \sqrt{K_a/C}$, $[H^+] = \sqrt{K_aC}$, $pH = \frac{1}{2}(pK_a - \log C)$; valid for $\alpha < 5\%$; α rises on dilution

WEAK BASE

$[OH^-] = \sqrt{K_bC}$

$pOH = \frac{1}{2}(pK_b - \log C)$, $pH = 14 - pOH$

WHY IT WORKS

Dilution favours ionisation (Le Chatelier on $HA \rightleftharpoons H^+ + A^-$): α rises as C falls, but $[H^+] = C\alpha = \sqrt{K_aC}$ still falls. K_a itself does not change with dilution, only with temperature.

POLYPROTIC ACIDS AND COMMON ION EFFECT

- ✓ H_3PO_4 and H_2CO_3 ionise in steps with $K_{a1} \gg K_{a2} \gg K_{a3}$; nearly all H^+ comes from the first step, $[A^{2-}] \approx K_{a2}$ for a pure H_2A solution.
- ✓ **Common ion effect**: adding a salt with an ion common to a weak electrolyte pushes the equilibrium back and lowers its ionisation (CH_3COONa added to CH_3COOH lowers $[H^+]$).

QUICK CHECK

pH and α of 0.1 M CH_3COOH ($K_a = 1.8 \times 10^{-5}$)?

Answer: $[H^+] = \sqrt{1.8 \times 10^{-6}} = 1.34 \times 10^{-3}$; pH $= 2.87$; $\alpha = 1.34\%$.

QUICK CHECK

pH of 0.1 M NH_3 ($K_b = 1.8 \times 10^{-5}$)?

Answer: $[OH^-] = 1.34 \times 10^{-3}$, pOH 2.87, pH $= 11.13$.

5 Salt hydrolysis, buffers and titration

A **salt** dissolved in water can react with it (**hydrolysis**) if it contains the conjugate of a weak acid or weak base. The ion of a strong acid or strong base does not react. So the pH of a salt solution depends on which parent is weak.

PH OF SALT SOLUTIONS

SALT OF	EXAMPLE	HYDROLYSIS	PH FORMULA
strong acid + strong base	$NaCl$, KNO_3	none	7
weak acid + strong base	CH_3COONa, Na_2CO_3	$A^- + H_2O \rightleftharpoons HA + OH^-$ (basic)	$pH = 7 + \frac{1}{2}pK_a + \frac{1}{2}\log C$
strong acid + weak base	NH_4Cl, $CuSO_4$	$BH^+ + H_2O \rightleftharpoons B + H_3O^+$ (acidic)	$pH = 7 - \frac{1}{2}pK_b - \frac{1}{2}\log C$
weak acid + weak base	CH_3COONH_4	both ions hydrolyse	$pH = 7 + \frac{1}{2}(pK_a - pK_b)$, independent of C

HYDROLYSIS CONSTANT

$K_h = \dfrac{K_w}{K_a}$ (anion), $\dfrac{K_w}{K_b}$ (cation)

$h = \sqrt{K_h/C}$ degree of hydrolysis (weak acid + weak base: $h = \sqrt{K_w/K_aK_b}$)

ACIDIC BUFFER

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

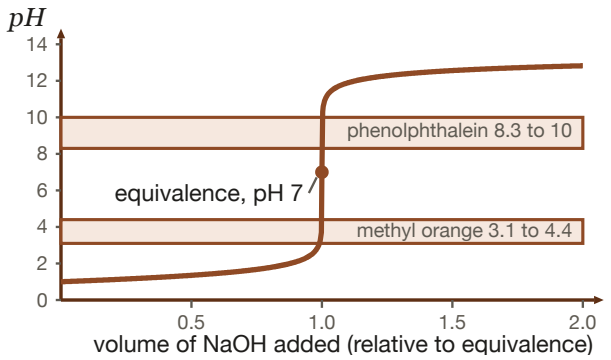
Henderson; CH_3COOH + CH_3COONa, $pK_a = 4.74$

BASIC BUFFER

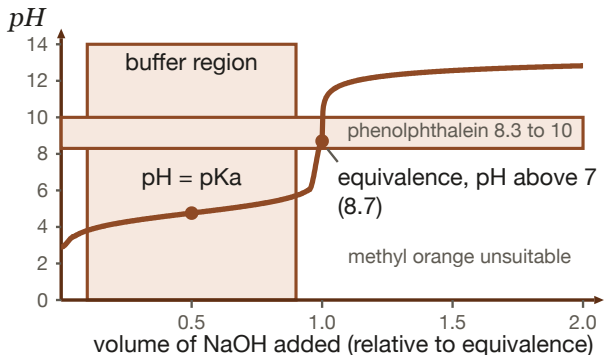
$$pOH = pK_b + \log\frac{[\text{salt}]}{[\text{base}]}$$

NH_3 + NH_4Cl, $pK_b = 4.74$

BUFFER A solution that resists a change in pH on adding small amounts of acid or base: a weak acid with its conjugate base (or a weak base with its conjugate acid) in comparable amounts. The acid component takes up added OH^-, the base component takes up added H^+. Buffer capacity is greatest when $[\text{salt}] = [\text{acid}]$, where $pH = pK_a$ (effective range $pK_a \pm 1$).



steep jump through pH 4 to 10, so both indicators work



half-equivalence gives $\text{pH} = \text{p}K_a$; salt hydrolysis puts the end point above 7

TITRATION AND INDICATORS

TITRATION	EQUIVALENCE-POINT PH	SUITABLE INDICATOR
strong acid + strong base	7	methyl orange (3.1 to 4.4) or phenolphthalein (8.3 to 10); the jump spans pH 4 to 10
weak acid + strong base	> 7 (hydrolysis of $\mathrm{A^-}$)	phenolphthalein; at half-equivalence $pH = pK_a$
strong acid + weak base	< 7	methyl orange
weak acid + weak base	near 7 but no sharp jump	no indicator works well

COMMON INDICATORS

INDICATOR	PH RANGE	COLOUR CHANGE
methyl orange	3.1 to 4.4	red to yellow
methyl red	4.4 to 6.2	red to yellow
litmus	5 to 8	red to blue
phenol red	6.8 to 8.4	yellow to red
phenolphthalein	8.3 to 10	colourless to pink

BUFFER FACTS

- A buffer is good only within $pK_a \pm 1$ (ratio $[\text{salt}]/[\text{acid}]$ between 0.1 and 10).
- Adding x mol $\mathrm{OH^-}$ turns x mol acid into salt: $pH = pK_a + \log \dfrac{s + x}{a - x}$; adding acid does the reverse.
- Dilution does not change the pH of a buffer (the ratio is unchanged) but lowers its capacity.
- Blood is buffered by $\mathrm{H_2CO_3/HCO_3^-}$ near pH 7.4; $\mathrm{CH_3COOH/CH_3COO^-}$ (4.74) and $\mathrm{NH_3/NH_4^+}$ (9.26) are the textbook pairs.

TRAP

Do not use Henderson once the weak acid is fully neutralised: that is a salt solution (hydrolysis). After mixing an acid and a base, first do the stoichiometry in mmol, then classify what is left: excess strong acid or base, a buffer, or a salt only. 10^{-8} M HCl is pH 6.98, never 8.

QUICK CHECK

100 mL of 0.1 M $\mathrm{CH_3COOH}$ is mixed with 50 mL of 0.1 M NaOH ($pK_a = 4.74$). pH?

Answer: 5 mmol acid left and 5 mmol $\mathrm{CH_3COO^-}$ formed: half-neutralised, $pH = pK_a = 4.74$.

6 Solubility product

SOLUBILITY PRODUCT For a sparingly soluble salt $\mathrm{A_xB_y\,(s) \rightleftharpoons} \, x\,\mathrm{A^{y+}\,(aq)} + y\,\mathrm{B^{x-}\,(aq)}$ in a saturated solution, $K_{sp} = [\mathrm{A^{y+}}]^x[\mathrm{B^{x-}}]^y$ (at fixed T). If the molar solubility is s: $K_{sp} = x^x y^y\, s^{x+y}$. A solid forms (precipitation) when the ionic product $Q_{sp} > K_{sp}$; the solid dissolves when $Q_{sp} < K_{sp}$.

K_{sp} IN TERMS OF THE MOLAR SOLUBILITY s

TYPE	EXAMPLES	K_{sp}	s
AB	AgCl, $\mathrm{BaSO_4}$	s^2	$\sqrt{K_{sp}}$
$\mathrm{AB_2}$ or $\mathrm{A_2B}$	$\mathrm{CaF_2}$, $\mathrm{Ag_2CrO_4}$, $\mathrm{Ca(OH)_2}$	$4s^3$	$(K_{sp}/4)^{1/3}$
$\mathrm{AB_3}$ or $\mathrm{A_3B}$	$\mathrm{Al(OH)_3}$, $\mathrm{Fe(OH)_3}$	$27s^4$	$(K_{sp}/27)^{1/4}$
$\mathrm{A_3B_2}$ or $\mathrm{A_2B_3}$	$\mathrm{Ca_3(PO_4)_2}$	$108s^5$	$(K_{sp}/108)^{1/5}$
$\mathrm{A_3B_4}$ (JEE favourite)	$\mathrm{Zr_3(PO_4)_4}$	$6912\,s^7$	$(K_{sp}/6912)^{1/7}$

COMMON ION

$s \approx$
$\frac{K_{sp}}{C}$ (AB), $\frac{K_{sp}}{C^2}$ (AB_2 in C of B)

neglect s beside C; solubility falls

PRECIPITATION TEST

$Q_{sp} > K_{sp}$

mix, dilute (volumes add), then compute ionic product; the first to cross K_{sp} precipitates first

PH AND SOLUBILITY

$M(OH)_n$: $[OH^-] = ns$

buffering the pH fixes $[OH^-]$ and then $s = K_{sp}/[OH^-]^n$; salts of weak acids dissolve more in acid

TRAP

Compare solubilities from K_{sp} only for salts of the **same type**. AgCl ($K_{sp} = 1.8 \times 10^{-10}$, $s = 1.3 \times 10^{-5}$ M) and Ag_2CrO_4 ($K_{sp} = 1.1 \times 10^{-12}$, $s = 6.5 \times 10^{-5}$ M): the smaller K_{sp} is the **more** soluble salt because the powers differ. Never forget $x^x y^y$, and that volumes add on mixing.

QUICK CHECK

K_{sp} of AgCl is 1.8×10^{-10}. Solubility in water and in 0.1 M NaCl?

Answer: Water: $\sqrt{1.8 \times 10^{-10}} = 1.34 \times 10^{-5}$ M. In NaCl: $s \approx \frac{1.8 \times 10^{-10}}{0.1} = 1.8 \times 10^{-9}$ M.

7 Worked examples

DISSOCIATION FROM PRESSURE

1 mol of N_2O_4 is 75% dissociated at total pressure P. K_p?

$N_2O_4 \rightleftharpoons 2\,NO_2$: $K_p = \frac{4\alpha^2 P}{1-\alpha^2} = \frac{4(0.5625)P}{0.4375} = 5.14\,P$

VAPOUR DENSITY

PCl_5 (vapour density 104.25 undissociated) has $d = 69.5$ at equilibrium. α?

$\alpha = \frac{D-d}{(n-1)d} = \frac{104.25 - 69.5}{69.5} = 0.50$

COMBINING CONSTANTS

$A \rightleftharpoons B$ has $K_1 = 4$ and $B \rightleftharpoons C$ has $K_2 = 0.5$. K for $C \rightleftharpoons A$?

$A \rightleftharpoons C$: $K = K_1 K_2 = 2$; reverse: $K = 1/2 = 0.5$

DIRECTION OF REACTION

For $H_2 + I_2 \rightleftharpoons 2\,HI$, $K_c = 50$. A mixture has $[H_2] = [I_2] = 0.1$, $[HI] = 0.8$ M. Which way?

$Q = \frac{0.8^2}{0.1 \times 0.1} = 64 > 50$: backward

PH OF A SALT

pH of 0.1 M CH_3COONa ($pK_a = 4.74$) and of 0.02 M NH_4Cl ($K_b = 10^{-5}$)?

$7 + \frac{1}{2}(4.74) + \frac{1}{2}\log 0.1 = 8.87$; $7 - \frac{1}{2}(5) - \frac{1}{2}\log 0.02 = 5.35$

AMMONIUM ACETATE

pH of CH_3COONH_4 with $K_a = K_b = 1.8 \times 10^{-5}$?

$7 + \frac{1}{2}(pK_a - pK_b) = 7$ (neutral, but the ions are hydrolysed)

BUFFER

pH of 0.1 M CH_3COOH and 0.2 M CH_3COONa ($pK_a = 4.74$)?

$pH = 4.74 + \log 2 = 5.04$

MIXING ACID AND BASE

50 mL of 0.1 M HCl and 50 mL of 0.2 M NaOH. pH?

excess OH^-: $10 - 5 = 5$ mmol in 100 mL $= 0.05$ M; pOH 1.30, pH 12.70

SOLUBILITY PRODUCT

$Ca(OH)_2$ has $K_{sp} = 5.5 \times 10^{-6}$. Solubility and pH of the saturated solution?

$s = (K_{sp}/4)^{1/3} = 0.0111$ M; $[OH^-] = 0.0222$, pOH 1.65, pH 12.35

PRECIPITATION

Equal volumes of 2×10^{-5} M Ag^+ and 2×10^{-5} M Cl^- are mixed. Does $AgCl$ ($K_{sp} = 1.8 \times 10^{-10}$) precipitate?

volumes double, so each is 10^{-5} M: $Q = 10^{-10} < 1.8 \times 10^{-10}$: no precipitate

EQUILIBRIUM MOLES

2 mol A and 2 mol B react in a 1 L flask: $A + B \rightleftharpoons C + D$, $K_c = 4$. Equilibrium amounts?

$K_c = \dfrac{x^2}{(2-x)^2} = 4 \Rightarrow \dfrac{x}{2-x} = 2$, $x = \dfrac{4}{3}$: C = D = 1.33 mol, A = B = 0.67 mol

COMMON ION EFFECT

α of 0.1 M CH_3COOH ($K_a = 1.8 \times 10^{-5}$) after adding 0.1 M CH_3COONa?

$[H^+] = K_a \dfrac{[\text{acid}]}{[\text{salt}]} = 1.8 \times 10^{-5}$ M, so $\alpha = \dfrac{1.8 \times 10^{-5}}{0.1} = 0.018\%$ (down from 1.34%); pH $= 4.74$

MIXING STRONG ACIDS

10 mL of 0.1 M HCl is mixed with 30 mL of 0.05 M HNO_3. pH?

$[H^+] = \dfrac{1.0 + 1.5}{40} = 0.0625$ M; pH $= 1.20$

HYDROLYSIS

K_h and degree of hydrolysis of 0.1 M CH_3COONa ($K_a = 1.8 \times 10^{-5}$)?

$K_h = \dfrac{10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$; $h = \sqrt{K_h/C} = 7.5 \times 10^{-5}$

8 Last-minute checklist

BEFORE YOU ANSWER

Which constant (K_c, K_p, K_a, K_{sp})? Is the species a solid (leave it out)? Strong or weak? Do the stoichiometry of any mixing before you use a formula. Did the volumes add?

- $K = \dfrac{[C]^c[D]^d}{[A]^a[B]^b}$; $K_p = K_c(RT)^{\Delta n_g}$; reverse $1/K$; $\times n$: K^n; add: multiply; $Q < K$ forward; only T changes K.
- PCl_5: $K_p = \alpha^2 P/(1-\alpha^2)$; N_2O_4: $4\alpha^2 P/(1-\alpha^2)$; $\alpha = (D-d)/((n-1)d)$; $\Delta G^\circ = -2.303RT \log K$.
- Le Chatelier: more reactant forward; pressure up towards fewer moles; inert gas at constant V no effect; heating favours the endothermic direction; catalyst no shift.
- $K_w = 10^{-14}$; $pH + pOH = 14$; $K_a K_b = K_w$; $[H^+] = \sqrt{K_a C}$; $pH = \frac{1}{2}(pK_a - \log C)$; very dilute acid: add 10^{-7}.
- Salts: WA+SB $pH = 7 + \frac{1}{2}pK_a + \frac{1}{2}\log C$; SA+WB $pH = 7 - \frac{1}{2}pK_b - \frac{1}{2}\log C$; WA+WB $pH = 7 + \frac{1}{2}(pK_a - pK_b)$; buffer $pH = pK_a + \log(\text{salt}/\text{acid})$.
- $K_{sp} = x^x y^y s^{x+y}$: AB s^2, AB_2 $4s^3$, AB_3 $27s^4$, A_3B_4 $6912s^7$; precipitate if $Q_{sp} > K_{sp}$; common ion cuts solubility.

Now test it. Practise this chapter with past-paper and bank questions on Prep Vault. **Practice this chapter →**