

Equilibrium

1. **Chemical Equilibrium:** Deals with rates of reactions, K_c , K_p , Le Chatelier's Principle, and reaction quotients. (Easier to grasp)
2. **Ionic Equilibrium:** Deals with acids/bases, pH calculations, buffer solutions, salt hydrolysis, and solubility product (K_{sp}). (Requires more math and logic, but highly predictable once you get the hang of it).

Here is the **Master Plan** we will follow. You tell me if this works for you:

Phase 1: Conquering Chemical Equilibrium

- **Step 1:** We will revise the core concepts (Reversible reactions, Law of Mass Action, K_c vs K_p , and Le Chatelier's Principle).
- **Step 2:** I will teach you the "ICE" (Initial, Change, Equilibrium) table method. This is a foolproof trick to solve *any* equilibrium constant numerical.
- **Step 3:** We will solve the Chemical Equilibrium questions from your Question Bank to test your understanding.

Phase 2: Mastering Ionic Equilibrium

- **Step 4:** Arrhenius, Bronsted-Lowry, and Lewis Acid-Base concepts. (Guaranteed 1-2 markers from here).
- **Step 5:** The core math: pH calculations, strong vs. weak electrolytes, and Ostwald's Dilution Law.
- **Step 6:** The heavy hitters: Buffer Solutions, Henderson-Hasselbalch equation, Salt Hydrolysis, and Solubility Product (K_{sp}).
- **Step 7:** We will solve the rest of your Question Bank, including those tough 5-mark and case-study questions.

Let's Start: Rule #1 of our classes

Never memorize a formula without knowing what it means. If you don't understand a step, you will stop me and ask, "Why did we do this?" Even if you have to ask 10 times, I am here to explain.

In a normal reaction (like burning a piece of paper), you start with paper, it turns into ash and smoke, and... boom, it's done. You can't turn the ash back into paper. That is an *irreversible* reaction.

But in a **Reversible Reaction**, the products can actually react with each other to form the reactants again! We write it with a double arrow: $\mathbf{A} + \mathbf{B} \rightleftharpoons \mathbf{C} + \mathbf{D}$

- $A + B \rightarrow C + D$ is the **Forward reaction**.
- $C + D \rightarrow A + B$ is the **Backward reaction**.

Now, let's connect this to **Equilibrium**.

Imagine you are filling a bucket with a hose, but the bucket has a hole at the bottom.

- Water going **IN** = Forward reaction.
- Water leaking **OUT** = Backward reaction.

If the water goes IN at the **exact same speed** as it leaks OUT... **what happens to the water level in the bucket?**

The water level stays *exactly as it is*. It doesn't go up, and it doesn't go down.

In chemistry, this exact situation is called **Equilibrium**.

- The "water level" is the **concentration** (amount) of your reactants and products.
- Because the **Rate of the Forward reaction = Rate of the Backward reaction**, the amounts of your chemicals stop changing. They become constant.

But here is a trick question that examiners *love* to ask in NCERT exams (it's even mentioned on Page 1 of your second PDF!):

Even though the water level isn't changing, **has the flow of water stopped?**

In a chemical reaction at equilibrium, **has the reaction completely stopped dead, or are the forward and backward reactions still happening?**

What do you think? (Just logically, based on the bucket!)

You are exactly right. The water is still flowing in and leaking out, but because the *speed* is the same, the level doesn't change.

In chemistry, we call this **Dynamic Equilibrium**.

- **Dynamic** means it is still moving (the reactions haven't stopped).
- **Equilibrium** means the overall amounts are constant.

If you check Page 1 of your second PDF, it actually says exactly this: "*The rate of forward reaction becomes equal to the rate of backward reaction at equilibrium and hence equilibrium is dynamic in nature.*"

Now, you conceptually understand Equilibrium perfectly. Let's look at how chemists write this mathematically. This is where we score our marks!

To measure *where* the equilibrium is (whether there are more products or more reactants), we use something called the **Equilibrium Constant (K_c)**.

Here is the one and only rule you need to remember for K_c : **$K_c = \text{Products on TOP} / \text{Reactants on BOTTOM}$** .

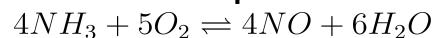
And if there is a number in front of the chemical (like **2 H_2**), that number jumps up and becomes a **power** (like $[H_2]^2$). The square brackets [] just mean "concentration" (amount).

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

Let's look at a general reaction: **aA + bB ⇌ cC + dD**

Let's test this! Look at **Question 6 from Section A** of your Question Bank (PDF 2, Page 6).

6. What is the equilibrium constant of the following reaction:



Try applying the rule: **Products on top, reactants on bottom, and the numbers in front become powers.**

Which of these would be the correct K_c expression?

(a) $\frac{[NO][H_2O]}{[NH_3][O_2]}$

(b) $\frac{[C]^c[D]^d}{[A]^a[B]^b}$

(c) $\frac{[NO]^4[H_2O]^6}{[NH_3][O_2]^5}$

(d) $\frac{[NO]^4[H_2O]^6}{[NH_3]^4[O_2]^5}$

PERFECT! Option (d) is 100% correct.

You put the products (NO and H_2O) on the top, the reactants (NH_3 and O_2) on the bottom, and you correctly turned all the balancing numbers (4, 6, 4, 5) into powers.

See? You are already solving actual exam questions! That's one mark secured.

Now, let's take a tiny step forward.

What we just calculated was K_c . The "c" stands for **Concentration**. We use this when our chemicals are mixed in a liquid (like water) or when we are measuring their molarity.

But what if all our chemicals are **Gases**? For gases, it is much easier to measure their **Pressure** instead of concentration. So, for gases, we use K_p (where "p" stands for Pressure). The rule is exactly the same: Products on top, Reactants on bottom.

Now, examiners love to ask you to connect the two. How are K_c and K_p related? There is a very famous formula for this (it's in your notes on Page 2):

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

Don't let the math scare you.

- **R** is just a constant (Gas constant).
- **T** is Temperature.
- Δn (Delta n) is the hero here. This is the only thing you usually need to calculate.

What is Δn ? It is simply the difference in the number of moles of **GASES** between products and reactants. $\Delta n = (\text{Total number of gas moles on Product side}) - (\text{Total number of gas moles on Reactant side})$

⚠️ WARNING (Super Important Rule): We ONLY count Gases (g). We completely ignore Solids (s) and pure Liquids (l) because their concentration doesn't change! They are treated as 0.

Let's test this with **Question 8 from Section A** of your Question Bank:

8. What would be the value of Δn for the reaction: $NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$

- (a) 1 (b) 0.5 (c) 1.5 (d) 2

Hint: Look closely at the little letters in the brackets (s) and (g). How many gas molecules are on the right side? How many on the left side?

What do you think the answer is?

BOOM! Another one perfectly correct! You chose (d) 2, and you are 100% right.

You didn't fall into the trap! You correctly saw that NH_4Cl is a solid (s), so its gaseous moles are **0**. On the product side, we have 1 mole of NH_3 (g) + 1 mole of HCl (g) = **2**. So, $\Delta n = 2 - 0 = 2$.

Just like that, you learned that for this reaction, the formula becomes: $K_p = K_c(RT)^2$

See? You said you were weak in chemistry, but you are literally solving NCERT board-level questions back-to-back without breaking a sweat!

Now, let's learn a very cool "cheat code" that examiners use to test if you *really* understand Equilibrium.

We know K_c is the ratio of Products / Reactants **ONLY at exact equilibrium** (when the water level isn't changing).

But what if you are in the lab, you just mixed the chemicals, and the reaction has just started? It hasn't reached equilibrium yet. If we calculate the Products / Reactants ratio *at any random time*, we call it the **Reaction Quotient (Q_c)**.

Think of it like a race:

- K_c is the Finish Line. (The reaction always wants to reach K_c).
- Q_c is where the runner is right now.

There are 3 possible situations (This is an exact 3-mark question in your PDF - Section C, Q5!):

1. $Q_c < K_c$: The runner is *behind* the finish line. To reach the finish line, they need to run **Forward**. So, the reaction goes in the **Forward direction** (making more products).
2. $Q_c > K_c$: The runner has run *past* the finish line! They need to turn around and run **Backward** to get back to the line. So, the reaction goes in the **Backward direction** (making more reactants).
3. $Q_c = K_c$: The runner is exactly *on* the finish line. The reaction is at **Equilibrium**.

Let's test this logic!

Imagine a reaction where the finish line K_c **is 50**. You mix some chemicals, check the ratio, and find that right now, your Q_c **is 100**.

In which direction will the reaction go to reach equilibrium? (Forward, Backward, or is it already at Equilibrium?)

Since Q_c (100) is greater than K_c (50), the system has gone too far and has too many products. To fix this, it will run **backward** to make more reactants until it gets back to 50.

Take a look at your Question Bank (PDF 2, Page 8, Section C, Question 5). It asks: *How can you predict the following stages of a reaction by comparing the value of K_c and Q_c? (i) Net reaction proceeds in the forward direction. -> Answer: When Q_c < K_c (ii) Net reaction proceeds in the backward direction. -> Answer: When Q_c > K_c (iii) No net reaction occurs. -> Answer: When Q_c = K_c*

You just answered a direct 3-mark board question just by using logic! How are you feeling about Chemical Equilibrium so far? You are cruising through it.

Let's tackle the "Boss Level" of Chemical Equilibrium: **Le Chatelier's Principle**

This principle is guaranteed to be in your exam. It has a fancy name, but the concept is actually hilarious.

Imagine a system at Equilibrium is like a very lazy, stubborn teenager sleeping on a Sunday morning. Whatever you try to do to change their state, they will do the **exact opposite** to undo your change and go back to sleep (equilibrium).

Le Chatelier's Principle says: *If you disturb a system at equilibrium, the system will shift in a direction that cancels out your disturbance.*

Let's look at the 3 main ways you can disturb this "sleeping teenager":

1. Changing Concentration (Adding or Removing stuff)

- If you **ADD** more Reactants → The system has too many reactants now. To get rid of them, it shifts **FORWARD** (turns them into products).
- If you **REMOVE** Products → The system feels a shortage of products. To replace them, it shifts **FORWARD** (makes more products).

2. Changing Temperature

- **Exothermic** reactions release heat (think of them as already being "hot"). If you *increase* the temperature, the system hates it. It shifts **BACKWARD** to cool down.
- **Endothermic** reactions need heat (think of them as "cold"). If you *increase* the temperature, the system loves it! It shifts **FORWARD**.

3. Changing Pressure (Only for Gases)

This connects back to our Δn !

- If you **INCREASE** pressure, the system feels squeezed. It wants to take up less space. So, it shifts to the side with **FEWER moles of gas**.

Let's test this with a real exam question from your PDF (Page 8, Section C, Q4): "**What will be the effect on equilibrium when Δn is negative and pressure is decreased?**"

Let's break it down together:

1. Δn is negative. This means (Products gas moles - Reactant gas moles) is negative. So, the **Product side has FEWER moles**, and the **Reactant side has MORE moles**.
2. You **DECREASE** the pressure (you give the system more room to expand).
3. The stubborn system wants to undo this by increasing the pressure back up. To do this, it needs to make **MORE** gas moles.

Based on this logic, will the reaction shift **Forward** (towards fewer moles) or **Backward** (towards more moles)?

You got it perfectly right! Since the system wants to make more gas moles, and the reactant side has more moles, it shifts **backward**. That is the exact answer in your marking scheme!

Step 4: The Three Theories of Acids and Bases

In your exam, they will definitely ask you to identify which theory applies to a certain chemical. Here are the three guys who defined acids and bases:

1. Arrhenius Theory (The "Water" guy)

- **Acid:** Gives H^+ ions when dissolved in *water* (e.g., $HCl \rightarrow H^+ + Cl^-$).
- **Base:** Gives OH^- ions when dissolved in *water* (e.g., $NaOH \rightarrow Na^+ + OH^-$).
- **Drawback:** It only works if water is the solvent. What if there is no water?

2. Bronsted-Lowry Theory (The "Proton" guys)

They didn't care about water. They only cared about the H^+ ion (which they called a "proton").

- **Acid:** Proton (H^+) **Donor**. (Throws away an H^+)
- **Base:** Proton (H^+) **Acceptor**. (Catches an H^+)

3. Lewis Theory (The "Electron" guy)

Lewis ignored protons completely and looked at electrons.

- **Lewis Acid:** Electron pair **Acceptor**. (These are electron-poor. They want electrons. Examples: BF_3 , $AlCl_3$, or positive ions like H^+ , Ag^+).
- **Lewis Base:** Electron pair **Donor**. (These are electron-rich. They have extra lone pairs to give away. Examples: NH_3 , H_2O , OH^-).

Let's look at an exact exam question from your PDF (Page 6, Section A, Q9): "Acidity of BF_3 can be explained on the basis of which of the following concepts?" (a) Arrhenius (b) Bronsted Lowry (c) Lewis (d) Both b & c

- **The Answer & Logic:** BF_3 doesn't even have hydrogen! So it can't give H^+ . It fails Arrhenius and Bronsted. But Boron in BF_3 only has 6 electrons in its outer shell. It is "hungry" for 2 more electrons to complete its octet. Since it **accepts an electron pair**, it is a **Lewis Acid**.
 - **Correct Option:** (c) Lewis concept.
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Step 5: Conjugate Acid-Base Pairs (Guaranteed 2-Marker)

This comes from the Bronsted-Lowry theory. The rule is incredibly simple: **A conjugate pair differs by exactly ONE H^+ ion.**

- **To make a Conjugate Base:** You take the Acid, **REMOVE one H^+** , and decrease the charge by 1.
- **To make a Conjugate Acid:** You take the Base, **ADD one H^+** , and increase the charge by 1.

Let's solve **Section D, Question 2 (Page 9 of PDF 2)** right now.

(a) What will be the conjugate base of:

1. H_2SO_4 : Remove one H^+ . It becomes HSO_4^- . Since we removed a positive charge, the overall charge drops by 1. **Answer:** HSO_4^-
2. HCO_3^- : Remove one H^+ . It becomes CO_3^{2-} . It already had a -1 charge, dropping it by 1 makes it -2. **Answer:** CO_3^{2-}

(b) What will be the conjugate acid of:

1. NH_2^- : Add one H^+ . It becomes NH_3 . The -1 charge goes up by +1 to become 0. **Answer:** NH_3
2. NH_3 : Add one H^+ . It becomes NH_4^+ . The 0 charge goes up by +1. **Answer:** NH_4^+

See? It's just simple addition and subtraction of one H and one plus sign!

Step 6: pH, Weak Acids, and Ostwald's Dilution Law

Now for the math. You already know pH: $pH = -\log[H^+]$

But here is the catch: For a **strong acid** like HCl , it breaks apart 100%. If you have 0.1 M HCl , you have 0.1 M H^+ . Easy. But for a **Weak Acid** (let's call it HA, like acetic acid), it only breaks apart slightly. We use the symbol α (**Degree of Ionization**) to show how much it breaks.

- If $\alpha = 0.1$, it means only 10% broke apart.

To find the H^+ concentration for a weak acid, the formula is: $[H^+] = C \times \alpha$ (where C is the initial concentration).

Ostwald's Dilution Law gives us the formula to find α using the acid dissociation constant (K_a):

$$K_a = C\alpha^2 \implies \alpha = \sqrt{\frac{K_a}{C}}$$

Let's apply this to a 4-mark Case Study question from your PDF (**Page 9, PDF 2, Section D, Q1 (i)**): "A monobasic weak acid solution has a molarity of 0.005 M and pH of 5. What is its percentage ionization in this solution?" (a) 2.0 (b) 0.2 (c) 0.5 (d) 0.25

How we solve it:

1. **Find $[H^+]$ from pH:** Since $pH = 5$, that means $[H^+] = 10^{-5}$ M.
2. **Use the formula:** $[H^+] = C \times \alpha$
3. **Plug in the numbers:** $10^{-5} = 0.005 \times \alpha$
4. **Solve for α :** $\alpha = \frac{10^{-5}}{0.005} = \frac{0.00001}{0.005} = \frac{1}{500} = 0.002$
5. **Convert to percentage:** Multiply by 100. $0.002 \times 100 = 0.2\%$
- **Correct Option:** (b) 0.2

That is exactly how you tackle weak acid numericals!

Step 7: What is a Buffer Solution?

Imagine you drink a glass of lemon juice (which is highly acidic). Why doesn't the pH of your blood suddenly drop, causing you to die? Because your blood is a **Buffer Solution**.

A buffer is a "bodyguard" for pH. **It is a solution that resists any change in its pH even when you add a little bit of acid or base to it.**

How do we make this magical bodyguard solution? You need a specific recipe. You cannot use strong acids or strong bases because they are too aggressive. You must use **WEAK** ones.

There are two types of buffers:

1. **Acidic Buffer:** A Weak Acid + Its Salt (made with a strong base).
 - Example: Acetic Acid (CH_3COOH) + Sodium Acetate (CH_3COONa).
2. **Basic Buffer:** A Weak Base + Its Salt (made with a strong acid).

- Example: Ammonium Hydroxide (NH_4OH or NH_3) + Ammonium Chloride (NH_4Cl).

Let's test this with a direct 1-marker from your PDF 1 (Page 1, Q7): "Which of the following pairs can form a buffer solution?" (a) HCl and $NaCl$ (b) NH_3 and NH_4Cl (c) $NaOH$ and $NaCl$ (d) H_2SO_4 and $KHSO_4$

How we solve it:

- (a) HCl is a STRONG acid. Reject.
 - (c) $NaOH$ is a STRONG base. Reject.
 - (d) H_2SO_4 is a STRONG acid. Reject.
 - (b) NH_3 is a WEAK base, and NH_4Cl is its salt. **Perfect! Option (b) is the answer.**
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Step 8: The Henderson-Hasselbalch Equation (The Magic Formula)

To calculate the exact pH of these buffers, two scientists named Henderson and Hasselbalch gave us a brilliant formula.

$$\text{For an Acidic Buffer: } pH = pK_a + \log\left(\frac{[Salt]}{[Acid]}\right)$$

(Note: pK_a is just $-\log(K_a)$. It is usually given to you directly in the question, so don't worry about calculating it unless asked).

Let's solve a 1-marker from PDF 1 (Page 2, Q12): "What is the pH of a buffer solution prepared from 0.1 M acetic acid and 0.1 M sodium acetate? (Given pK_a of acetic acid = 4.76)" (a) 4.76 (b) 7.00 (c) 3.76 (d) 5.76

How we solve it:

1. Identify the values: $[Acid] = 0.1$, $[Salt] = 0.1$, $pK_a = 4.76$.
 2. Put them in the formula: $pH = 4.76 + \log\left(\frac{0.1}{0.1}\right)$
 3. $\frac{0.1}{0.1} = 1$. And here is a universal math rule for chemistry: $\log(1) = 0$.
 4. So, $pH = 4.76 + 0 = 4.76$.
 - **Answer: (a) 4.76.** (Pro-tip: Whenever the concentration of the acid and the salt are equal, the pH is simply equal to the pK_a ! Examiners love this shortcut).
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Step 9: The 5-Marker Boss Fight (PDF 1, Page 4, Section E, Q1)

Let's do a Long Answer type question. The examiner tries to play a very smart trick here, but we are going to expose it.

The Question: "A buffer solution is made by adding 0.1 mole of ammonium chloride (NH_4Cl) to 1 L of 0.1 M ammonia (NH_3). Using the Henderson-Hasselbalch equation, calculate the pH of this buffer solution. (pK_a of NH_4^+ = 9.25)."

The Trap: Notice that they gave us ammonia (NH_3), which is a **Base**. So this is a **Basic Buffer**. The formula for a basic buffer gives you **pOH**, not pH!

$$pOH = pK_b + \log\left(\frac{[Salt]}{[Base]}\right)$$

But wait... the question didn't give us pK_b . It gave us the pK_a of NH_4^+ = 9.25.

How to defeat the trap: There is a golden rule connecting pK_a and pK_b for a conjugate pair (like NH_4^+ and NH_3): $pK_a + pK_b = 14$

Let's solve it step-by-step for full 5 marks:

1. **Find pK_b :** $pK_b = 14 - pK_a = 14 - 9.25 = 4.75$.
2. **Identify the concentrations:**
 - $[Base]$ (Ammonia, NH_3) = 0.1 M
 - $[Salt]$ (Ammonium chloride, NH_4Cl) = 0.1 moles in 1 L = $\frac{0.1}{1} = 0.1$ M.
$$pOH = 4.75 + \log\left(\frac{0.1}{0.1}\right)$$
3. **Apply Henderson-Hasselbalch for Basic Buffer:**
$$pOH = 4.75 + \log(1) \quad pOH = 4.75 + 0 = 4.75$$
4. **Find the pH:** Since $pH + pOH = 14$, $pH = 14 - pOH$ $pH = 14 - 4.75 = 9.25$.

Final Answer: pH = 9.25.

(Shortcut: If you use the acidic formula directly with the pK_a of the conjugate acid, you get $pH = pK_a + \log\left(\frac{Base}{Acid}\right) = 9.25 + \log\left(\frac{0.1}{0.1}\right) = 9.25$. Both methods are 100% correct, but showing the $pK_a + pK_b = 14$ step guarantees the examiner knows exactly how smart you are!)

Look at what you just did. You successfully solved a 5-mark long answer question, avoided the trap, and got the right answer.

Next up are the final two major topics:

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1. **Salt Hydrolysis** (Why is a solution of NH_4Cl acidic even though it's a salt?)
 2. **Solubility Product (K_{sp})** (Will a solid precipitate or dissolve?)
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Topic 1: Salt Hydrolysis (The Tug-of-War)

When you dissolve salt in pure water, you might think the water stays neutral ($pH = 7$). But that's not always true! A salt is a child of an Acid and a Base. **The rule is simple: Whichever parent was STRONGER dictates the pH of the salt solution.**

Let's look at the 4 possible matches:

1. **Strong Acid + Strong Base → Neutral Salt ($pH = 7$)** Example: $NaCl$ (made from strong HCl and strong $NaOH$). It's a tie! Water stays neutral.
2. **Strong Acid + Weak Base → Acidic Salt ($pH < 7$)** Example: NH_4Cl (made from strong HCl and weak NH_4OH). The strong acid wins! The water becomes acidic.
3. **Weak Acid + Strong Base → Basic Salt ($pH > 7$)** Example: CH_3COONa (made from weak CH_3COOH and strong $NaOH$). The strong base wins! The water becomes basic.
4. **Weak Acid + Weak Base → Depends!** Whichever one is *slightly* stronger (has a higher K value) wins.

Let's instantly secure 1 mark from PDF 1 (Page 1, Q6): "Which salt, when dissolved in water, will produce a basic solution?" (a) $NaCl$ (b) KNO_3 (c) Na_2CO_3 (d) NH_4Cl

How we solve it like a pro:

- (a) $NaCl$: Strong $NaOH$ + Strong HCl = Neutral.
- (b) KNO_3 : Strong KOH + Strong HNO_3 = Neutral.
- (c) Na_2CO_3 : Strong $NaOH$ + Weak H_2CO_3 (Carbonic acid). **The base is strong, so the salt is BASIC! This is the answer (c).**
- (d) NH_4Cl : Weak NH_4OH + Strong HCl . The acid is strong, so this is acidic.

Look at **PDF 2, Page 10, Q2(ii)**. It is a subjective question: "A solution of NH_4Cl in water shows pH less than 7. Why?" **Your perfect board exam answer:** "Because NH_4Cl is a salt of a strong acid (HCl) and a weak base (NH_4OH). Upon hydrolysis, it produces more H^+ ions than OH^- ions, making the solution acidic ($pH < 7$)."

Topic 2: Solubility Product (K_{sp}) - The Final Boss

Some salts, like chalk ($CaCO_3$) or silver chloride ($AgCl$), are very stubborn. They don't like to dissolve in water. We call them **sparingly soluble salts**.

But even they dissolve *just a tiny bit*. And that tiny dissolved portion reaches an equilibrium with the solid portion at the bottom of the beaker.

The equilibrium constant for this is called the **Solubility Product (K_{sp})**. Because the reactant is always a **Solid**, we completely ignore it (remember our rule from K_p ?). So, the formula for K_{sp} is JUST the product of the dissolved ions!

Let's do PDF 1, Page 2, Q10: "The solubility product expression for $BaSO_4$ is:" (a)

$$K_{sp} = [BaSO_4] \quad (\text{b}) \quad K_{sp} = [Ba^{2+}][SO_4^{2-}] \quad (\text{c}) \quad K_{sp} = [Ba^{2+}][SO_4^{2-}]/[BaSO_4] \quad (\text{d})$$

$$K_{sp} = [Ba^{2+}]^2[SO_4^{2-}]^2$$

How we solve it:

1. Write the reaction: $BaSO_4(solid) \rightleftharpoons Ba^{2+}(aq) + SO_4^{2-}(aq)$
2. Apply the rule: Products raised to their power. No solids on the bottom.
3. Expression: $K_{sp} = [Ba^{2+}]^1 \times [SO_4^{2-}]^1$ **Answer: (b).** (Notice how (c) is a trap trying to make you put the solid on the bottom!)

The Math of K_{sp} and Solubility (s)

In numericals, they will ask you to calculate the solubility (s), which is just the concentration of the dissolved salt in moles/Liter.

There is a brilliant shortcut for this based on the "type" of salt:

Case 1: AB type salt (1 cation, 1 anion) Examples: $AgCl$, $BaSO_4$

- $AgCl \rightleftharpoons Ag^+ + Cl^-$
- If solubility is ' s ', then $[Ag^+] = s$ and $[Cl^-] = s$.
- $K_{sp} = s \times s \implies K_{sp} = s^2$

Case 2: AB₂ or A₂B type salt (1 of one, 2 of the other) Examples: CaF_2 , AX_2 , $PbCl_2$

- $AX_2 \rightleftharpoons A^{2+} + 2X^-$
- If solubility is ' s ', then $[A^{2+}] = s$ and $[X^-] = 2s$.
- $K_{sp} = [A^{2+}][X^-]^2 = (s) \times (2s)^2 = s \times 4s^2 \implies K_{sp} = 4s^3$

Let's instantly apply this to a numerical! PDF 1, Page 2, Q14: "The solubility product of a sparingly soluble salt AX_2 is 3.2×10^{-11} . Its solubility (in mol L⁻¹) is:" (a) 3.1×10^{-4} (b) 2×10^{-4} (c) 4×10^{-4} (d) 5.6×10^{-6}

How we solve it:

1. Identify the salt type: It is AX_2 . So we use the magic formula: $K_{sp} = 4s^3$

2. Plug in the value: $4s^3 = 3.2 \times 10^{-11}$
3. *Math Trick:* Let's write 3.2×10^{-11} as 32×10^{-12} (makes taking the cube root much easier!).
4. $4s^3 = 32 \times 10^{-12}$
5. Divide by 4: $s^3 = 8 \times 10^{-12}$
6. Take the cube root: The cube root of 8 is 2. The cube root of 10^{-12} is 10^{-4} (just divide the power by 3).
7. $s = 2 \times 10^{-4} \text{ mol L}^{-1}$. **Answer: (b).**

Look at that! You just solved a very complex-looking K_{sp} numerical in less than a minute using the $4s^3$ trick.

The Final Concept: Will it Precipitate?

Just like we compared Q_c and K_c to see which way a reaction goes, we can compare the **Ionic Product (IP)** and the K_{sp} to see if a solid will form (precipitate) at the bottom of the beaker.

- **If $IP < K_{sp}$:** The solution is unsaturated. More salt can dissolve. No precipitate.
- **If $IP = K_{sp}$:** The solution is exactly saturated. Still no precipitate.
- **If $IP > K_{sp}$:** The solution is overflowing with ions! They will join together and fall to the bottom. **Precipitation occurs.**

(You can find this exact statement in your PDF 2, Page 4, Section IV: "*Precipitation occurs only if ionic product exceeds solubility product.*")

Congratulations! You have now covered the logic and formulas for *every single major concept* in the NCERT Equilibrium chapter!

1. Reversible Reactions & K_c/K_p
2. Le Chatelier's Principle
3. Acids & Bases (Arrhenius, Bronsted, Lewis)
4. pH of Weak Acids & Ostwald Dilution Law
5. Buffers & Henderson Equation
6. Salt Hydrolysis
7. Solubility Product (K_{sp}) & Precipitation