

Equilibrium & Acid/Base

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What is Equilibrium?

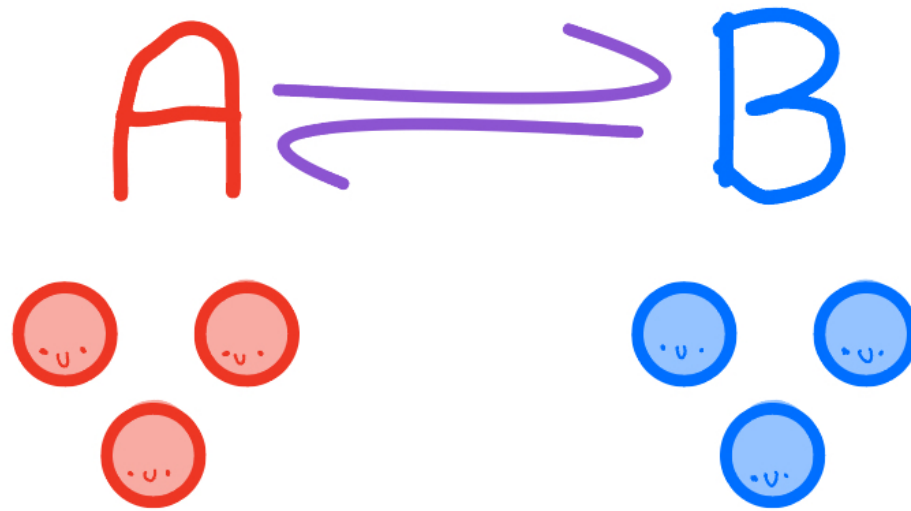
Equilibrium refers to the idea of DYNAMIC BALANCE - think “wildly mediocre” : whole lot of nothing, but also everything.

- Essentially, there is an equal rate of forward and reverse reaction

This is the ideal state of a chemical reaction. Baking-soda vinegar reactions exhibit equilibrium once there are no more bubbles or fizzing - there is an equal concentration of reactants and products.

In the human body, equilibrium is seen in all homeostatic processes: hemoglobin and oxygen; blood-clotting; pH of blood (will be covered in Acid Base section)

System at Equilibrium



Equal concentrations, equal
rate of forward and reverse
reaction

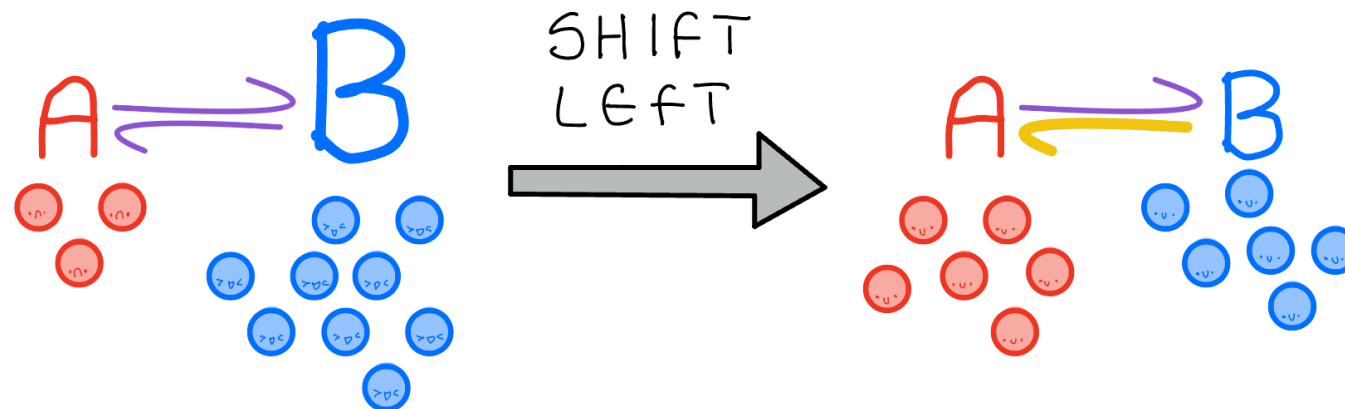
Le Chatelier's Principle

States that when a system that is in a state of equilibrium is subjected to a change in concentration, pressure, or temperature, the system will shift to counteract the change and return to equilibrium. Almost all questions are based around changes in concentration.

To counteract the change the system experiences, it will do something to change the rate of reaction, causing a shift to the left or to the right. In simpler terms, a shift right results in an increase of products, a shift left results in an increase of reactants.

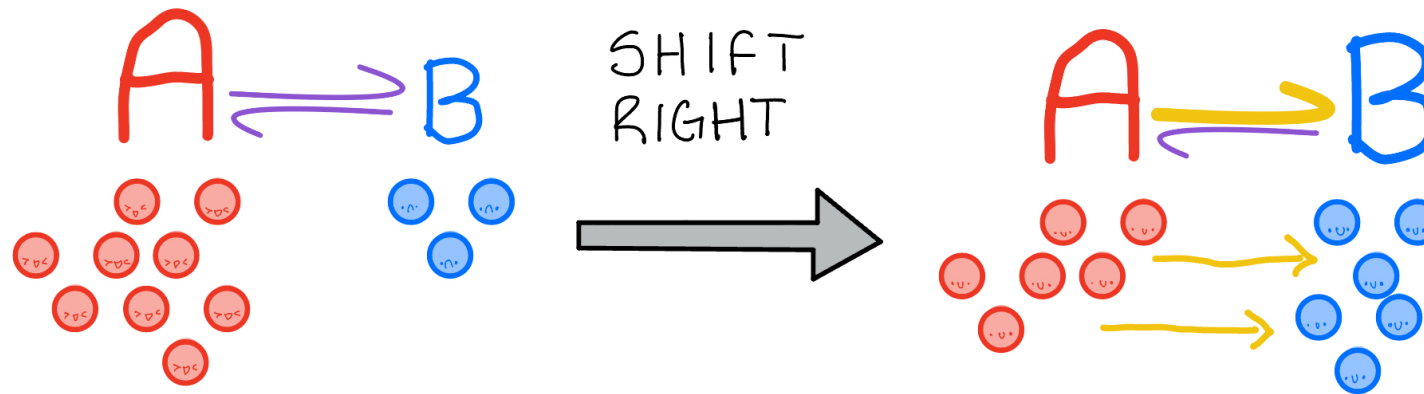
Let's look at some examples...

Increase: Products



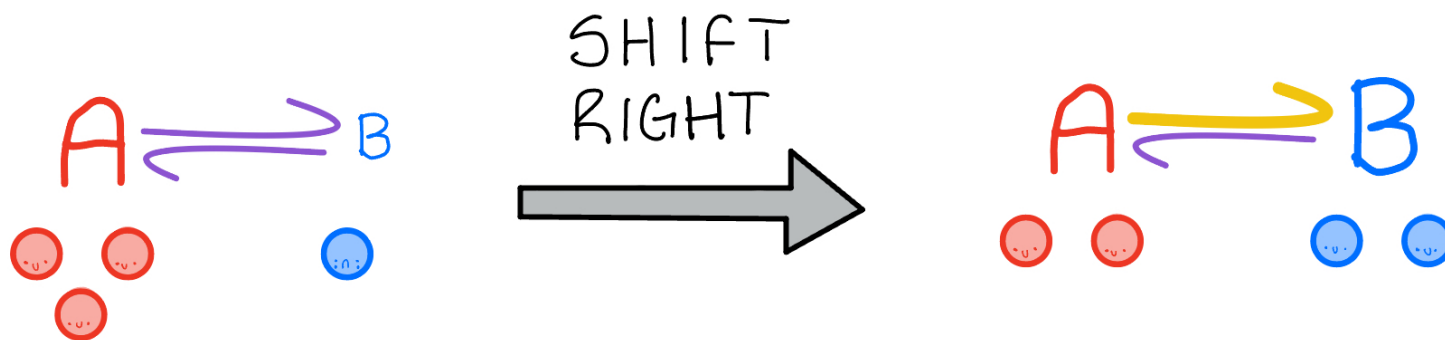
Compensation: increase rate of **reverse** reaction to create more reactants; shift system **left**

Increase: Reactants



Compensation: increase rate of **forward** reaction to create more products; shift system **right**

Decrease: Products



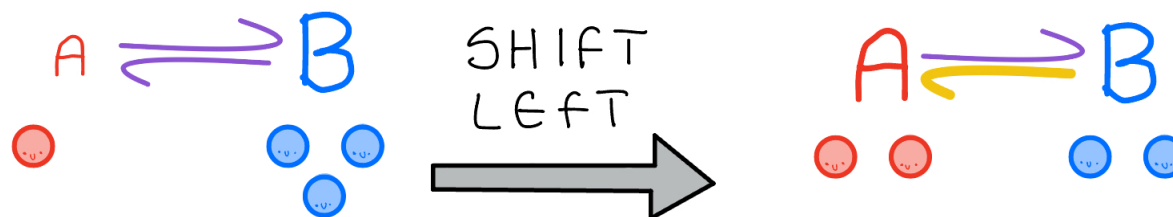
Compensation: increase rate of **forward** reaction to create more products; shift system **right**

Decrease: Reactants - Your Turn



What will the compensation be?

Decrease: Reactants - Answer



What will the compensation be? Shift
left

Medical Relevance

The two core examples of equilibrium in the human body are:

1. Hemoglobin and Oxygen

- a. When in the lungs, there is a high concentration (partial pressure of Oxygen), which causes hemoglobin to shift to favour oxygen binding and CO₂ release, creating oxyhemoglobin
- b. When in the tissues, there is the inverse, causing hemoglobin to release oxygen and take up CO₂.

2. Acid Base Balance in the Blood

- a. More detail later on - try to think about how pH is affected by concentration, and what compensations we will see

Equilibrium Constant

Equilibrium constant is the math behind a system's state of equilibrium; specifically, the amount of products and reactants needed to achieve equilibrium. Note that equilibrium constant can **only** be calculated for gaseous and aqueous substances - NO SOLIDS because they eventually dissolve to (aq) forms. If (s) is written, don't use it!

The value, K_c , represents the ratio of reactants to products needed to attain equilibrium. If the value is higher than 1, there are more products; if the value is less than one, there are more reactants.

$$K_c = \frac{[C]^x [D]^x}{[A]^x [B]^x}$$

Ratio of products
to reactants at
equilibrium

$K_c > 1$ = more product
 $K_c < 1$ = more reactant
 $K_c = 1$ = equal amounts

$$K_c = \frac{[C]^x [D]^x}{[A]^x [B]^x}$$

Coefficient of
product
[Product]

Coefficient of
reactant
[Reactant]

Example 1: $2\text{HI}(\text{g}) \rightleftharpoons \text{H}_2(\text{g}) + \text{I}_2(\text{g})$

$$K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2}$$

Example 1: $7\text{HCl(g)} \rightleftharpoons \text{H}_2\text{(g)} + 3\text{Cl}_2\text{(g)}$

Your turn - note that this reaction is not possible, we're just practicing writing!

Example 1: $7\text{HCl(g)} \rightleftharpoons \text{H}_2\text{(g)} + 3\text{Cl}_2\text{(g)}$

$$K_c = \frac{[\text{H}_2][\text{Cl}_2]^3}{[\text{HCl}]^7}$$

Your turn - note that this reaction is not possible, we're just practicing writing!

Example 2: $4\text{PO}_4(\text{aq}) + 4\text{C}(\text{s}) \rightleftharpoons 7\text{CO}_2(\text{g}) + 3\text{P}(\text{g})$

Your turn - watch out for (s)!

Example 2: $4\text{PO}_4(\text{aq}) + 4\text{C}(\text{s}) \rightleftharpoons 7\text{CO}_2(\text{g}) + 3\text{P}(\text{g})$

$$K_c = \frac{[\text{CO}_2]^7 [\text{P}]^3}{[\text{PO}_4]^4}$$

Your turn - watch out for (s)!

Using the Formula

Questions about equilibrium constant will either give you the value of K_c , or the amount of moles of each product and/or reactant. In the case of being given K_c , and amounts of only products/reactants, you will need to create an ICE table.

	CO (g)	+	H ₂ O	⇌	CO ₂ (g)	+	H ₂ (g)
Initial	0.100 M		0.100 M		0		0
Change	- x		-x		x		x
Equilibrium	0.100 -x		0.100 -x		x		x

The 'E' values will be put into the formula, with K_c already being known. Then you must solve for x .

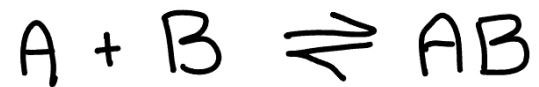
Using the Formula

Exams that cover as much as Midterm 1 does will not make you do all this work. You will be given the initial values of all products and reactants. Let's go through a question together!

“For the reaction $A + B \rightarrow AB$ at a temperature of 25°C at equilibrium, the concentrations of A, B and AB are 0.01M. Calculate the equilibrium constant K.”

Since no values about states are given, assume everything is (g) or (aq). Since there is no change in concentration, pressure, or temperature, there is no need to do extra calculations. Keeping in mind the formula and the values provided, try to solve this on your own. The answer is on the next slide - no peeking!

Write your answer here!



$$A = 0.01 \text{ M}$$

$$B = 0.01 \text{ M}$$

$$AB = 0.01 \text{ M}$$

$$K_c > 1$$

↳ MORE PRODUCTS

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

$$K_c = \frac{0.01}{(0.01)(0.01)}$$

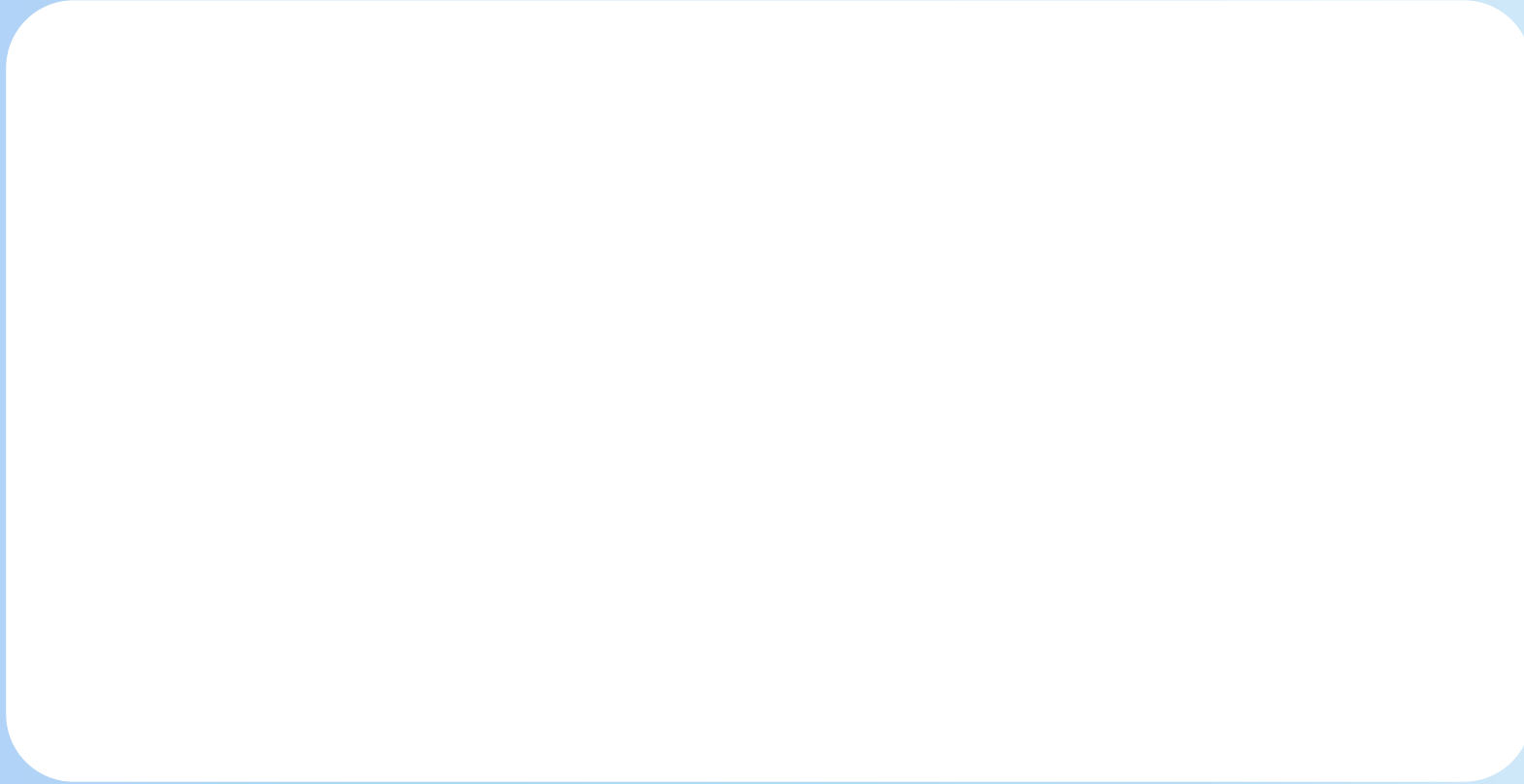
$$K_c = \frac{0.01}{0.0001} = 100$$

Using the Formula

Another one! - *DJ Khalid*

“For the reaction $2A(s) + B(g) \rightarrow 2C(g) + D(aq)$ where the concentration of A is 1 M, B is 2.5 M, C is 1 M, and D is 3 M, calculate the equilibrium constant K.”

Write your answer here!





$$A = 1 \text{ M}$$

$$B = 2.5 \text{ M}$$

$$C = 1 \text{ M}$$

$$D = 3 \text{ M}$$

$$K_c < 1$$

↳ MORE REACTANTS

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

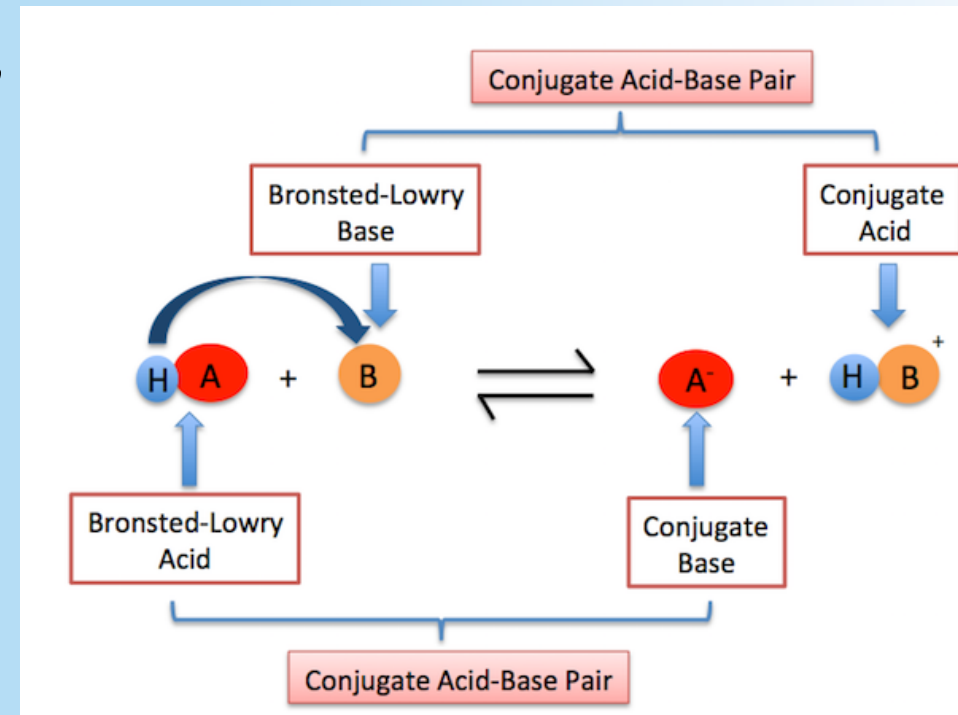
$$K_c = \frac{[1]^2 [2.5]}{[3]}$$

$$K_c = \frac{2.5}{3} = 0.\overline{83}$$

Acids and Bases - Brønsted-Lowry Theory

In the Brønsted-Lowry theory of acids and bases, acids are proton donors, and bases are proton acceptors. Protons in this theory are always hydrogen atoms (H^+).

When acids donate their proton, they become a **conjugate base**. When bases accept a proton, they become a **conjugate acid**. A reaction between an acid and base will always yield a conjugate acid and a conjugate base. Effectively, they're swapping roles!



Acids and Bases - Dissociation

Dissociation refers to what happens when acids and bases dissolve into smaller pieces - almost always ions. The strength of an acid/base affects how they dissociate.

The values K_a and K_b refer to the ability of an acid or base to donate or accept protons.



Acid Dissociation - Ka



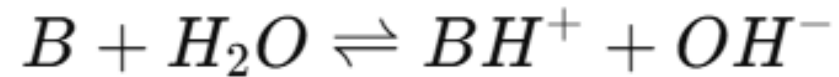
$$K_a = \frac{[H^+][A^-]}{[HA]}$$

Large Ka = strong acid
Small Ka = weak acid

More common in biochem
Small pKa = strong acid
(dissociates easily)
Large pKa = weak acid
(holds on to H⁺)

$$pK_a = -\log K_a$$

Base Dissociation - K_b



$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

Large K_a = strong acid
Small K_a = weak acid

$$pK_b = -\log K_b$$

More common in biochem
Small pK_b = strong base
Large pK_b = weak base

Important Formulas

$$pH + pOH = 14$$

pH of strong acid

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

pOH of strong base

$$pOH = -\log [OH^-]$$

pH of weak acid

$$pH = -\log \sqrt{K_a \times C_a}$$

pOH of weak base

$$pOH = -\log \sqrt{K_b \times C_b}$$

Refers to
conjugate acid

Practice Question

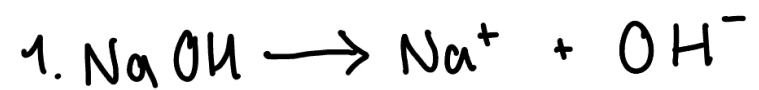
“The pH of a 0.001 M NaOH solution is equal to:

- A. 13
- B. 11
- C. 7
- D. 4
- E. 3

Write your answer here!



Solution



2. $[\text{OH}^-] = 0.001 \text{ M}$

3. $\text{pOH} = -\log [\text{OH}^-]$
 $= -\log [0.001] = 3$

4. $\text{pH} + \text{pOH} = 14$

$\text{pH} = 14 - \text{pOH}$

$\text{pH} = 14 - 3$

$\text{pH} = 11$

Practice Question

“The pH of a 0.001 M NaOH solution is equal to:

- A.
- B.
- C.
- D.
- E.

13

11

7

4

3

Buffers!

Buffers act as pH bodyguards that help to resist drastic changes in pH. A buffer consists of two parts: a **weak** acid/base & a **conjugate** acid/base.

You can think of buffers as a team:

- If too much acid is added -> the base “soaks up” extra H^+
- If too much base is added -> the acid gives H^+



Considering how useful these buffers are, it's no surprise that our body uses many physiological buffer systems! The most important one to know is the Bicarbonate Buffer System.

The Henderson-Hasselbalch Equation

pH > pKa = more basic
pH < pKa = more acidic
pH = pKa = equal []

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

conjugate
base
concentration

Buffers are most
effective within
+/- 1 pH unit of
pKa!!!

pKa = the dissociation
constant of the weak
acid

acid concentration

Bicarbonate Buffer System

Imbalance occurs when buffering capacity of bicarbonate is exceeded, causing a change in blood pH. Disorders will either have a respiratory or metabolic cause.



H₂CO₃ and HCO₃⁻ act as buffers to keep blood within its ideal pH, 7.4.

Blood pH is managed by both the lungs and kidneys.

The lungs maintain blood pH by adjusting levels of CO₂

- inc CO₂ -> dec pH
- dec CO₂ -> inc pH

The kidneys maintain blood pH by adjusting HCO₃⁻ levels

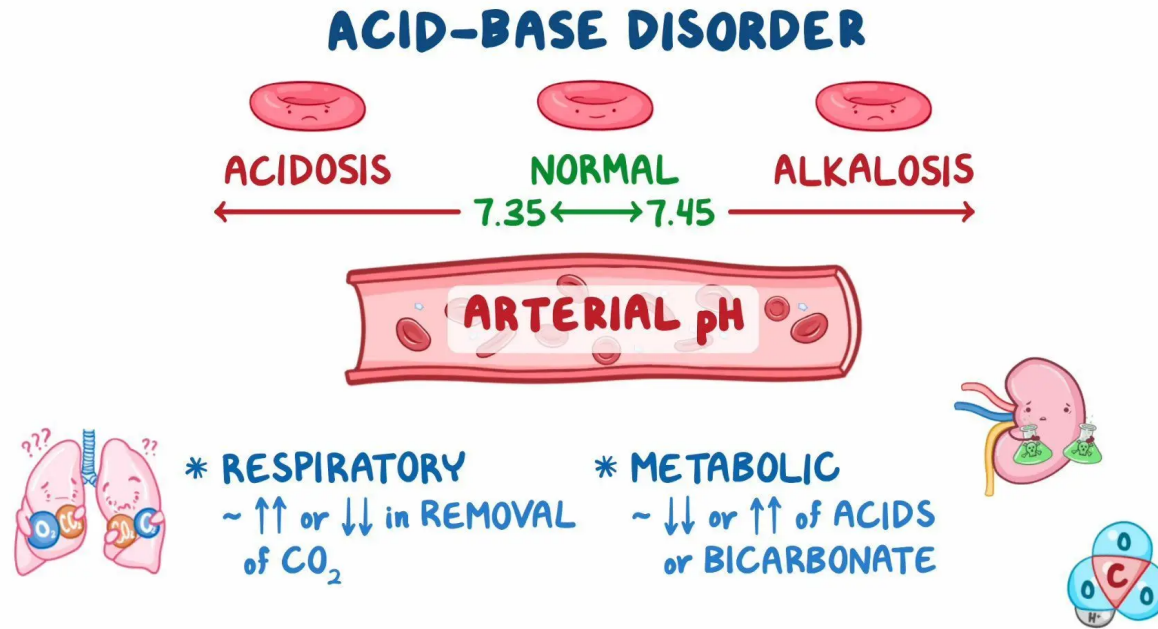
- inc HCO₃⁻ -> inc pH
- dec HCO₃⁻ -> dec pH

Acid Base Disorders

Disorder	Cause	Effect	Compensation
Metabolic acidosis	↑ acid (e.g., DKA, renal failure)	↓ HCO_3^-	Hyperventilation (↓ CO_2)
Metabolic alkalosis	↑ HCO_3^- (vomiting, diuretics)	↑ pH	Hypoventilation (↑ CO_2)
Respiratory acidosis	CO_2 retention (COPD)	↑ H^+	Kidneys ↑ HCO_3^- reabsorption
Respiratory alkalosis	Hyperventilation	↓ CO_2	Kidneys excrete more HCO_3^-

DISORDER	pH	pCO ₂	HCO ₃ ⁻	COMPENSATION
RESPIRATORY ACIDOSIS	↓ < 7.35	↑ > 45 mmHg	↑ OR NORMAL	KIDNEYS RETAIN HCO ₃ ⁻ (INTO BLOOD)
RESPIRATORY ALKALOSIS	↑ > 7.45	↓ < 35 mmHg	↓ OR NORMAL	KIDNEYS EXCRETE HCO ₃ ⁻ (INTO URINE)
METABOLIC ACIDOSIS	↓ < 7.35	↓ OR NORMAL	↓ < 22 mEq/L	LUNGS HYPERVENTILATE (pCO ₂ ↓)
METABOLIC ALKALOSIS	↑ > 7.45	↑ OR NORMAL	↑ > 26 mEq/L	LUNGS HYPOVENTILATE (pCO ₂ ↑)

Acid Base Disorders



USMLE Question Challenge



As hurricane Melissa hits, mandem starts hyperventilating yelling “BOMBOCLATT!”

Which of the following arterial blood gas findings would be expected in mandem?

- A) $\downarrow$ pH, $\uparrow$ CO_2
- B) $\uparrow$ pH, $\downarrow$ CO_2
- C) $\downarrow$ pH, $\downarrow$ CO_2
- D) Normal pH, normal CO_2

USMLE Question Challenge



As hurricane Melissa hits, mandem starts hyperventilating yelling “BOMBOCLATT!”

Which of the following arterial blood gas findings would be expected in mandem?

- A) $\downarrow$ pH, $\uparrow$ CO_2
- B) $\uparrow$ pH, $\downarrow$ CO_2
- C) $\downarrow$ pH, $\downarrow$ CO_2
- D) Normal pH, normal CO_2

WooClap Questions

Please do not use these to cheat if you are attending the seminar currently! WooClap during the seminar is to test your understanding, not for you to be the best (on an anonymous form, no less)

The following slides are for you to reference later, or for you to use as anki cards, or to use if you weren't able to attend the seminar!

Good luck :)

The pOH of a solution is 3.5. What is the pH?

1

3.5

0%

0 

2

12.5

0%

0 

3

11.5

0%

0 

4

10.5

0%

0 

5

9.5

0%

0 

The pOH of a solution is 3.5. What is the pH?

①

3.5

0% 0 

②

12.5

0% 0 

③

11.5

0% 0 

④

10.5

0% 0 ✓

⑤

9.5

0% 0 

If the products in a system in equilibrium are INCREASED, which direct will the system shift?

1

Left

0% 0 

2

No change

0% 0 

3

Right

0% 0 

If the products in a system in equilibrium are INCREASED, which direct will the system shift?

1

Left

0% 0 ✓

2

No change

0% 0 👤

3

Right

0% 0 👤

In the Brønsted–Lowry theory, a base is defined as:

- ① A proton donor 0% 0 
- ② A proton acceptor 0% 0 
- ③ A hydrogen gas 0% 0 
- ④ water 0% 0 
- ⑤ is that guy's last name Norwegian 0% 0 

In the Brønsted–Lowry theory, a base is defined as:

1

A proton donor

0% 0 

2

A proton acceptor

0% 0 ✓

3

A hydrogen gas

0% 0 

4

water

0% 0 

5

is that guy's last name Norwegian

0% 0 

A buffer works best when

1

It contains a strong acid and a strong base

0%

0 

2

It contains only water

0%

0 

3

$\text{pH} = \text{pK}_a \pm 1$

0%

0 

4

$\text{pH} = 14$

0%

0 

A buffer works best when

1

It contains a strong acid and a strong base

0%

0 

2

It contains only water

0%

0 

3

$\text{pH} = \text{pK}_a \pm 1$

0%

0 ✓

4

$\text{pH} = 14$

0%

0 

A 0.004 M barium hydroxide ($\text{Ba}(\text{OH})_2$) solution is prepared. What is the pOH of this solution?

1

2.00

0%

0



2

2.10

0%

0



3

1.90

0%

0



4

1.70

0%

0



0:56

A 0.004 M barium hydroxide ($\text{Ba}(\text{OH})_2$) solution is prepared. What is the pOH of this solution?

①

2.00

0% 0 

②

2.10

0% 0 ✓

③

1.90

0% 0 

④

1.70

0% 0 

In respiratory alkalosis

- 1

pH inc, pCO₂ inc

0% 0 
- 2

pH dec, pCO₂ inc

0% 0 
- 3

pH inc, pCO₂ dec

0% 0 
- 4

pH inc, pO₂ dec

0% 0 

In respiratory alkalosis

1

pH inc, pCO₂ inc

0% 0 

2

pH dec, pCO₂ inc

0% 0 

3

pH inc, pCO₂ dec

0% 0 ✓

4

pH inc, pO₂ dec

0% 0 

In metabolic acidosis

- 1

pH inc, HCO₃ dec

0% 0 
- 2

pH dec, HCO₃ dec

0% 0 
- 3

pH dec, HCO₃ inc

0% 0 
- 4

pH inc, HCO₃ inc

0% 0 

In metabolic acidosis

- ① pH inc, HCO_3 dec 0% 0 
- ② pH dec, HCO_3 dec 0% 0 ✓
- ③ pH dec, HCO_3 inc 0% 0 
- ④ pH inc, HCO_3 inc 0% 0 