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Acids and Bases

Study Guide — Acids, Bases & pH

70 Pre-med/IB-style questions covering Arrhenius vs Brønsted–Lowry vs Lewis definitions, conjugate pairs, amphiprotic species, strong/weak vs concentrated/dilute, pH/pOH and Kw, buffers, titrations, salt hydrolysis, and common conceptual traps.

70 items — Study Guide with Answers

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1 **In Brønsted–Lowry theory, an acid is best defined as:**

- A A substance that increases [OH−] in water
- B **A proton (H^+) donor** ✓
- C An electron-pair donor
- D A substance with pH > 7
- E Any molecule containing hydrogen

▶ **Explanation:** A Brønsted–Lowry acid donates a proton (H^+). Increasing [OH−] describes an Arrhenius base, electron-pair donor describes a Lewis base, pH>7 is not a definition, and having hydrogen is not sufficient to be an acid.

2 **In Brønsted–Lowry theory, a base is best defined as:**

- A **A proton (H^+) acceptor** ✓
- B A proton (H^+) donor
- C A substance that always contains OH− in its formula
- D A substance with pH < 7
- E A substance that releases electrons into solution

▶ **Explanation:** A Brønsted–Lowry base accepts a proton. Bases do not need to contain OH− (e.g., NH3), and pH < 7 describes acids, not bases.

3 **Which statement matches the Arrhenius definition of an acid?**

- A A substance that donates an electron pair
- B A substance that accepts a proton

- C **A substance that increases the concentration of H3O^{+} in aqueous solution ✓**
- D A substance that always reacts with metals to produce hydrogen gas
- E A substance that is always molecular

▶ **Explanation:** Arrhenius acids increase [H3O^{+}] (or [H^{+}]) in water. Proton-accepting is Brønsted base; electron-pair donation is Lewis base. Not all acids produce H2 with metals, and acids can be ionic or molecular.

4 **Which statement correctly defines a Lewis acid?**

- A A proton donor
- B **An electron-pair acceptor ✓**
- C An electron-pair donor
- D A substance that increases [OH−] in water
- E A substance with a bitter taste

▶ **Explanation:** A Lewis acid accepts an electron pair (e.g., BF3). Proton donor is Brønsted acid; electron-pair donor is Lewis base; [OH−] increase is Arrhenius base; taste is not a scientific definition.

5 **In the reaction NH3 + H2O NH4^{+} + OH−, which species is the Brønsted–Lowry acid?**

- A NH3
- B **H2O ✓**
- C NH4^{+}
- D OH−
- E All species are acids

▶ **Explanation:** The Brønsted acid donates H^+. Water donates H^+ to NH3 to form $NH4^+$, so H2O acts as the acid here. NH3 is the base (accepts H^+).

6 **In the reaction NH3 + H2O $NH4^+$ + OH−, what is the conjugate acid of NH3?**

- A H2O
- B NH2−
- C **$NH4^+$** ✓
- D OH−
- E $H3O^+$

▶ **Explanation:** A conjugate acid is formed when a base gains H^+. NH3 gains H^+ to become $NH4^+$.

7 **Which species is amphiprotic (can act as BOTH a Brønsted acid and a Brønsted base)?**

- A Cl−
- B $NH4^+$
- C **HCO3−** ✓
- D Na^+
- E CH4

▶ **Explanation:** HCO3− can donate H^+ to become $CO3^{2}$− (acid behavior) or accept H^+ to become H2CO3 (base behavior). Cl− and Na^+ do not donate/accept H^+ in this sense; $NH4^+$ is mainly an acid; CH4 is extremely weak as an acid/base in water.

8 **Which of the following is commonly classified as a strong acid in water?**

A $CH3COOH$

B HF

C **HCl** ✓

D $H2CO3$

E $NH4^+$

▶ **Explanation:** HCl is a strong acid (essentially complete ionization in water). Acetic acid, HF, and carbonic acid are weak acids; $NH4^+$ is a weak acid (conjugate acid of NH3).

9 **Which of the following is commonly classified as a strong base in water (in typical high-school chemistry)?**

A NH3

B **NaOH** ✓

C CH3COO−

D H2O

E CO2

▶ **Explanation:** NaOH is a strong base (dissociates to give OH−). NH3 is a weak base; acetate is a weak base; water is amphiprotic; CO2 is an acidic oxide in water.

10 **Which statement correctly distinguishes a strong acid from a concentrated acid?**

A A strong acid has a low pH, while a concentrated acid has pH $= 0$

B A strong acid fully ionizes in water, while a concentrated acid simply has a high molar concentration ✓

C A strong acid always contains oxygen, while a concentrated acid does not

D A strong acid is always dangerous, while a concentrated acid is always safe

E Strong acids have higher molar mass than concentrated acids

▶ **Explanation:** Strength refers to degree of ionization (how completely it forms ions). Concentration refers to how much acid is dissolved per liter. A weak acid can be concentrated, and a strong acid can be dilute.

11 At 25°C, the pH of pure water is closest to:

A 0

B 7 ✓

C 14

D 1

E Depends only on the container size

▶ **Explanation:** At 25°C, pure water autoionizes so that $[H3O^{+}] = [OH-] = 1.0\times10\hat{}-7$ M, giving pH = 7. (At other temperatures pH of neutral water changes, but 25°C is the standard reference.)

12 A solution has $[H3O^{+}] = 1.0 \times 10\hat{}-3$ M. What is its pH (approximately)?

A 3 ✓

B 7

C 11

D −3

E 0.001

▶ **Explanation:** pH = $-\log10[H3O^{+}]$. For $1.0\times10\hat{}-3$, pH = 3. The pH is not the concentration itself.

13 **If pH = 2, the hydronium concentration is closest to:**

- A **$1 \times 10\hat{}-2$ M** ✓
- B $2 \times 10\hat{}-2$ M
- C $1 \times 10\hat{}-12$ M
- D 2 M
- E 1×10^{2} M

▶ **Explanation:** $[H3O^{+}] = 10\hat{}-pH$, so for pH 2: $[H3O^{+}] = 10\hat{}-2$ M.

14 **If [OH−] = $1.0 \times 10\hat{}-4$ M at 25°C, what is the pH (approximately)?**

- A 4
- B 7
- C **10** ✓
- D 14
- E 18

▶ **Explanation:** pOH = $-\log10[OH-]$ = 4, and at 25°C pH + pOH = 14, so pH = 10.

15 **Which statement about pH is correct?**

A A pH change of 1 unit means $[H3O^+]$ changes by a factor of 2

B **Lower pH means higher $[H3O^+]$** ✓

C pH is directly equal to $[H3O^+]$

D Neutral solutions always have pH 7 at any temperature

E pH can never be below 0

▶ **Explanation:** pH is a log scale: lower pH corresponds to higher $[H3O^+]$. A 1-unit pH change corresponds to a $10\times$ change in $[H3O^+]$. Neutral pH depends on temperature, and very concentrated acids can have negative pH.

16 **At 25°C, if $[H3O^+] = [OH—]$, the solution is:**

A Always acidic (pH $<$ 7)

B Always basic (pH $>$ 7)

C **Neutral (pH $=$ 7)** ✓

D Impossible because $[H3O^+]$ and [OH—] cannot be equal

E Neutral only if the solution contains NaCl

▶ **Explanation:** At 25°C, neutrality means $[H3O^+] = [OH—] = 1.0\times10\hat{}—7$ M, which corresponds to pH 7. Equality indicates neutrality (at that temperature).

17 **Which statement about polyprotic acids is correct?**

A **They can donate more than one proton per molecule** ✓

B They must be strong acids

C They always donate all protons completely in water

D They have exactly one conjugate base

E They cannot form buffers

▶ **Explanation:** Polyprotic acids (e.g., H2CO3, H3PO4) can donate multiple protons stepwise, often with decreasing strength at each step. They have multiple conjugate bases and can form buffer systems.

18 **For the acid H2CO3, what is the conjugate base after it donates ONE proton?**

A $CO3^{2}$—

B **HCO3—** ✓

C $H3CO3^{+}$

D CO2

E OH—

▶ **Explanation:** Removing one proton from H2CO3 gives HCO3—. $CO3^{2}$— would require losing two protons total.

19 **Which conjugate acid–base pair differs by exactly one proton (H^{+})?**

A HCl / Cl2

B H2O / O2—

C **$NH4^{+}$ / NH3** ✓

D H2CO3 / CO2

E NaOH / Na^{+}

▶ **Explanation:** Conjugate pairs differ by one H^+. $NH4^+$ and NH3 differ by one proton. The other pairs do not differ by exactly one proton.

20 **Which statement about H2SO4 in water is most accurate at an introductory level?**

- A Both protons dissociate completely with equal strength
- B **The first dissociation is strong; the second dissociation is weaker (partial)** ✓
- C H2SO4 is a weak acid because it has two protons
- D H2SO4 cannot donate protons because sulfate is stable
- E H2SO4 is a base because it contains oxygen

▶ **Explanation:** Sulfuric acid is strong for the first proton (essentially complete), while the second dissociation is weaker. Treating both steps as equally complete is a common oversimplification.

21 **In the reaction HCl + H2O → Cl— + $H3O^+$, which species is the Brønsted base?**

- A HCl
- B **H2O** ✓
- C Cl—
- D $H3O^+$
- E None; this is not an acid–base reaction

▶ **Explanation:** The Brønsted base accepts H^+. Water accepts a proton from HCl to form $H3O^+$, so H2O is the base.

22 **What is the conjugate base of $H3O^+$?**

- A **H2O ✓**
- B OH–
- C H2O2
- D O2
- E H^+

▶ **Explanation:** A conjugate base is formed by removing one H^+ from the acid. Removing H^+ from $H3O^+$ gives H2O.

23 **Which statement about Ka is correct?**

- A A larger Ka means a weaker acid
- B **A larger Ka means a stronger acid ✓**
- C Ka is only defined for bases
- D Ka increases when an acid is diluted because it ionizes more
- E Ka equals the pH of the solution

▶ **Explanation:** Ka measures the extent of acid dissociation: larger Ka indicates more ionization and thus a stronger acid. Ka is an equilibrium constant (at fixed temperature), so it does not change just because you dilute the solution.

24 **Which statement about pKa is correct?**

- A **Lower pKa means stronger acid ✓**
- B Higher pKa means stronger acid

C pKa is the same as pH

D pKa is only used for strong acids

E pKa increases when temperature decreases for all acids

▶ **Explanation:** pKa = −log10(Ka). Stronger acids have larger Ka, so they have smaller pKa. pKa is a property of the acid equilibrium, not the same as pH.

25 Which is the best explanation for why HF is a weak acid in water while HCl is a strong acid?

A **HF has a stronger H–F bond that is harder to break (less ionization)** ✓

B Fluorine is less electronegative than chlorine

C HF contains oxygen, which makes acids weak

D HCl is weak because it has a larger molar mass

E HF is strong in water but looks weak due to pH scale limits

▶ **Explanation:** Despite fluorine's high electronegativity, the H–F bond is very strong and HF does not ionize completely in water. HCl ionizes essentially completely.

26 Which trend is generally correct for the acidity of hydrogen halides in water (HX) down Group 17?

A HF > HCl > HBr > HI

B **HI > HBr > HCl > HF** ✓

C HCl > HI > HF > HBr

D All hydrogen halides have identical acidity

E Acidity depends only on electronegativity, so HF is strongest

▶ **Explanation:** Down the group, H–X bonds become weaker and larger anions stabilize charge better, so acidity increases: HI is strongest, HF weakest.

27 Which order best describes the strength of the oxyacids of chlorine (from strongest to weakest)?

- A $HClO > HClO_2 > HClO_3 > HClO_4$
- B **$HClO_4 > HClO_3 > HClO_2 > HClO$** ✓
- C $HClO_2 > HClO_4 > HClO > HClO_3$
- D All have the same strength because they all contain chlorine
- E Strength depends only on the number of hydrogens, so they are equal

▶ **Explanation:** For oxyacids with the same central atom, more oxygens generally means stronger acid due to stronger electron-withdrawing effect and better resonance stabilization of the conjugate base.

28 Which statement best describes why chloroacetic acid ($ClCH_2COOH$) is stronger than acetic acid (CH_3COOH)?

- A Chloroacetic acid has a higher molar mass
- B **Chlorine withdraws electron density, stabilizing the conjugate base** ✓
- C Chlorine donates electrons, destabilizing the conjugate base
- D Acetic acid cannot form hydrogen bonds, but chloroacetic acid can
- E Chloroacetic acid is strong because it fully dissociates like HCl

▶ **Explanation:** An electron-withdrawing group (like Cl) stabilizes the negative charge on the conjugate base (carboxylate), making proton loss easier and increasing acidity.

29 **What is the pH of 0.010 M HCl (assuming complete dissociation and ignoring water's contribution)?**

A 1

B **2** ✓

C 7

D 12

E 14

▶ **Explanation:** For a strong acid, $[H3O^{+}]$ 0.010 = 10ˆ−2, so pH = 2.

30 **What is the pH of 0.010 M NaOH (assuming complete dissociation) at 25°C?**

A 2

B 7

C 10

D **12** ✓

E 14

▶ **Explanation:** [OH−] = 0.010 = 10ˆ−2 → pOH = 2 → pH = 14 − 2 = 12.

31 **A solution has pH 5. Compared to a solution with pH 3, it has:**

A 100 times higher $[H3O^{+}]$

B 10 times higher $[H3O^{+}]$

C 10 times lower $[H3O^{+}]$

D **100 times lower [$H3O^+$]** ✓

E The same [$H3O^+$]

▶ **Explanation:** A 2-unit pH increase corresponds to a $10^2 = 100\times$ decrease in [$H3O^+$]. pH 5 is $100\times$ less acidic (in terms of [$H3O^+$]) than pH 3.

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32 **Which mixture would form a buffer solution?**

A HCl(aq) + NaCl(aq)

B NaOH(aq) + NaCl(aq)

C **CH3COOH(aq) + CH3COONa(aq)** ✓

D HNO3(aq) + KNO3(aq)

E Pure water + sugar

▶ **Explanation:** A buffer requires a weak acid and its conjugate base (or a weak base and its conjugate acid). CH3COOH (weak acid) and CH3COO– (from CH3COONa) form a buffer. Strong acid + its salt (A) does not buffer effectively.

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33 **Which description best matches how a buffer resists a decrease in pH when a small amount of acid is added?**

A The weak acid component reacts with the added acid

B **The conjugate base component neutralizes added $H3O^+$ (or H^+)** ✓

C The buffer prevents any reaction from happening

D The buffer makes pH stay exactly 7 always

E The buffer works because it is concentrated

▶ **Explanation:** When acid is added, the conjugate base (A−) consumes H^+ to form HA, limiting the pH drop. Buffers do not freeze pH at 7; they resist changes around a characteristic pH range.

34 **In the buffer system HA/A−, when a small amount of strong base is added, the primary reaction is:**

- A $A^- + OH^- \rightarrow HA$
- B **$HA + OH^- \rightarrow A^- + H_2O$ ✓**
- C $HA + H_2O \rightarrow H_3O^+ + A^-$ (goes to completion)
- D $A^- + H_2O \rightarrow HA + OH^-$ (goes to completion)
- E No reaction occurs in a buffer

▶ **Explanation:** Added OH^- is neutralized by the weak acid component HA, producing water and converting HA to its conjugate base A−. This limits the rise in pH.

35 **A buffer works best when:**

- A [HA] is much larger than [A−]
- B [A−] is much larger than [HA]
- C **[HA] and [A−] are comparable in magnitude ✓**
- D The buffer contains a strong acid and a strong base
- E The buffer is diluted as much as possible

▶ **Explanation:** Buffer capacity is highest when both the weak acid and conjugate base are present in significant amounts (similar concentrations). If one component is tiny, it gets used up quickly and buffering fails.

36 **For a weak acid buffer, which statement is true when [A−] = [HA]?**

- A pH = 7 always
- B **pH = pKa** ✓
- C pH = 14 − pKa
- D pH = 0
- E pH is undefined for buffers

▶ **Explanation:** From Henderson–Hasselbalch: pH = pKa + log([A−]/[HA]). If [A−]=[HA], log(1)=0, so pH = pKa.

37 **Which titration has an equivalence point at pH 7 (at 25°C)?**

- A **Strong acid + strong base** ✓
- B Weak acid + strong base
- C Strong acid + weak base
- D Weak acid + weak base
- E Any neutralization always ends at pH 7

▶ **Explanation:** Strong acid–strong base titrations produce a neutral salt and water at equivalence, giving pH 7. Weak/strong combinations shift equivalence above or below 7 due to salt hydrolysis.

38 **At equivalence in a weak acid–strong base titration, the pH is typically:**

- A Less than 7, because acids always dominate
- B Exactly 7, because moles acid = moles base

C **Greater than 7, because the conjugate base of the weak acid hydrolyzes to form $OH-$** ✓

D Exactly 14, because base was added

E Impossible to predict

▶ **Explanation:** At equivalence, the solution contains the salt of the weak acid: its conjugate base reacts with water to produce $OH-$, making the solution basic ($pH > 7$).

39 **At equivalence in a strong acid–weak base titration, the pH is typically:**

A Greater than 7

B Exactly 7

C **Less than 7, because the conjugate acid of the weak base hydrolyzes to form $H3O^+$** ✓

D Exactly 0

E Always 14

▶ **Explanation:** The salt contains the conjugate acid of a weak base (e.g., $NH4^+$), which reacts with water to produce $H3O^+$, so the equivalence solution is acidic ($pH < 7$).

40 **Which indicator is MOST suitable for a strong acid–weak base titration (equivalence point < 7)?**

A **An indicator that changes color around pH 3–4** ✓

B An indicator that changes color around pH 7

C An indicator that changes color around pH 9–10

D Any indicator works equally well

E No indicator can be used for titrations

▶ **Explanation:** For strong acid–weak base titrations, the steep pH jump occurs in the acidic region, so you choose an indicator whose transition range overlaps that region (e.g., methyl orange). Indicators changing near 9–10 would miss the endpoint.

41 Which indicator is MOST suitable for a weak acid–strong base titration (equivalence point > 7)?

- A An indicator that changes color around pH 3–4
- B An indicator that changes color around pH 7
- C **An indicator that changes color around pH 9–10** ✓
- D Any indicator works equally well
- E Indicators only work for strong acid–strong base titrations

▶ **Explanation:** Weak acid–strong base titrations have equivalence in the basic range, so an indicator that changes color around pH 8–10 (e.g., phenolphthalein) is appropriate.

42 Which salt solution is expected to be approximately neutral (pH 7) at 25°C?

- A NH4Cl
- B CH3COONa
- C **NaCl** ✓
- D Na2CO3
- E AlCl3

▶ **Explanation:** NaCl is formed from a strong acid (HCl) and a strong base (NaOH), so neither ion hydrolyzes significantly → neutral. NH4Cl is acidic ($NH4^{+}$ hydrolyzes); CH3COONa and Na2CO3 are basic (anions hydrolyze); Al^{3+} tends to acidify water.

43 **Which salt solution is expected to be acidic?**

- A NaNO3
- B KCl
- C **NH4Cl** ✓
- D CH3COONa
- E NaF

▶ **Explanation:** NH4Cl contains NH4^{+} (conjugate acid of weak base NH3), which hydrolyzes to produce H3O^{+} $\rightarrow$ acidic. Nitrate and chloride from strong acids are essentially neutral; acetate and fluoride are conjugate bases of weak acids and tend to be basic.

44 **Which salt solution is expected to be basic?**

- A NaCl
- B NH4NO3
- C **CH3COONa** ✓
- D KBr
- E HCl(aq)

▶ **Explanation:** CH3COONa contains acetate (CH3COO—), the conjugate base of a weak acid, which hydrolyzes to produce OH— $\rightarrow$ basic. NaCl and KBr are neutral salts (strong acid/strong base). NH4NO3 tends to be acidic due to NH4^{+}.

45 **Which oxide is amphoteric (can react with both acids and bases) at a basic high-school level?**

A CO_2

B SO_3

C CaO

D **Al_2O_3** ✓

E Na_2O

▶ **Explanation:** Al_2O_3 is amphoteric: it can react with acids (acting basic) and with strong bases (acting acidic, forming aluminates). CO_2 and SO_3 are acidic oxides; CaO and Na_2O are basic oxides.

46 **Which statement about neutralization is correct?**

A Neutralization always produces a solution with pH exactly 7

B Neutralization always produces CO_2 gas

C **A key net ionic step is often $H_3O^+ + OH^- \rightarrow 2H_2O$** ✓

D Neutralization can only happen between strong acids and strong bases

E Neutralization means removing all ions from solution

▶ **Explanation:** The essential acid–base neutralization step is H^+ (or H_3O^+) reacting with OH^- to form water. The final pH depends on whether the acid/base are strong or weak and on stoichiometry.

47 **Which solution is expected to conduct electricity BEST (at the same molar concentration)?**

A $CH_3COOH(aq)$

B $NH_3(aq)$

C **HCl(aq)** ✓

D Pure water

E Sugar solution

▶ **Explanation:** HCl is a strong electrolyte (fully ionizes), producing many ions and high conductivity. Acetic acid and ammonia are weak electrolytes, pure water has very low ion concentration, and sugar does not ionize.

48 Which relationship is true for a conjugate acid–base pair at 25°C?

A Ka + Kb = Kw

B **Ka × Kb = Kw** ✓

C Ka = Kb always

D Ka = 1/Kb

E Ka and Kb are unrelated

▶ **Explanation:** For a conjugate pair HA/A– in water, Ka(HA) × Kb(A–) = Kw (1.0×10^–14 at 25°C).

49 If Ka for an acid HA is very large, then Kb for its conjugate base A– is:

A Very large

B **Very small** ✓

C Exactly 1

D Equal to Ka

E Impossible to relate

▶ **Explanation:** Ka × Kb = Kw. If Ka is large, Kb must be small to keep the product equal to Kw. Strong acids have weak conjugate bases.

50 **Which conjugate base is the STRONGEST base (in water) among the following?**

- A Cl—
- B NO3—
- C **CH3COO—** ✓
- D ClO4—
- E I—

▶ **Explanation:** Cl—, NO3—, and ClO4— are conjugate bases of strong acids and are extremely weak bases. Acetate (CH3COO—) is the conjugate base of a weak acid (acetic acid), so it is significantly more basic than the others listed.

51 **You mix 50.0 mL of 0.10 M HCl with 50.0 mL of 0.10 M NaOH. Assuming ideal behavior and complete reaction, the resulting solution is closest to:**

- A pH 2 (acidic)
- B **pH 7 (neutral)** ✓
- C pH 12 (basic)
- D pH 1 (very acidic)
- E pH 9 (slightly basic)

▶ **Explanation:** Equal moles of strong acid and strong base neutralize completely: H^+ + OH— → H2O. The remaining solution contains NaCl (neutral salt) in water, so pH 7 at 25°C.

52 **You mix 10.0 mL of 0.10 M HCl with 5.0 mL of 0.10 M NaOH. What best describes the result (qualitatively)?**

- A Neutral, because any acid mixed with any base becomes neutral
- B **Acidic, because there are more moles of HCl than NaOH** ✓
- C Basic, because NaOH is a strong base
- D Neutral, because HCl is strong
- E Cannot determine without Ka values

▶ **Explanation:** Compare moles: HCl moles = 0.0100 L×0.10 = 0.0010 mol; NaOH moles = 0.0050 L×0.10 = 0.00050 mol. Acid is in excess, so the solution remains acidic.

53 **Which statement about the pH of a 0.10 M strong acid solution vs a 0.10 M weak acid solution is correct?**

- A They must have the same pH because concentration is the same
- B **The strong acid solution has lower pH because it produces more $H3O^+$ at the same concentration** ✓
- C The weak acid solution has lower pH because weak acids are 'more reactive'
- D Weak acids produce more $H3O^+$ because they do not dissociate fully
- E pH depends only on molar mass, not dissociation

▶ **Explanation:** At equal formal concentration, a strong acid ionizes much more, giving a higher $[H3O^+]$ and therefore a lower pH than a weak acid.

54 **Which statement is TRUE about adding water to a solution of strong acid (no reaction, just dilution)?**

- A The number of moles of $H3O^+$ increases

B **The number of moles of $H3O^+$ stays (approximately) the same, but $[H3O^+]$ decreases** ✓

C pH decreases because dilution makes acids stronger

D pH becomes exactly 7 regardless of how much acid was present

E Dilution changes Ka

▶ **Explanation:** Dilution increases volume, so concentration decreases. For a strong acid, the moles of acid (and thus moles of $H3O^+$ produced) stay essentially constant, but $[H3O^+]$ decreases so pH increases.

55 A student says: 'Weak acids are dangerous, strong acids are safe.' Which correction is most accurate?

A Correct: weak acids always have higher pH than strong acids at any concentration, so they cannot harm you

B **Incorrect: strength describes ionization, but safety depends on concentration, reactivity, and exposure; both can be hazardous** ✓

C Correct: weak acids never donate protons

D Incorrect: strong acids do not ionize at all

E Correct: strong acids are always dilute by definition

▶ **Explanation:** Acid 'strength' is about dissociation extent, not hazard. Concentrated weak acids can be very corrosive, and strong acids can also be dangerous even when dilute. Safety depends on more than Ka.

56 Which statement about buffer capacity is correct?

A A buffer can neutralize unlimited acid/base without changing pH

B **A buffer has a limited capacity; once HA or A— is used up, pH can change rapidly** ✓

C A buffer works only if its pH is exactly 7

D Buffers only resist changes when base is added, not acid

E Buffers work by preventing any equilibrium shifts

▶ **Explanation:** Buffers resist pH change by consuming added H^+ or OH–, but only as long as significant amounts of the weak acid/base pair remain. Once one component is depleted, buffering breaks down.

57 Which pair is a major physiological buffer system often taught at high-school level?

A HCl / Cl–

B **H2CO3 / HCO3–** ✓

C NaOH / Na^+

D H2SO4 / $SO4^{2}$–

E HNO3 / NO3–

▶ **Explanation:** The carbonic acid/bicarbonate system (H2CO3/HCO3–) is a key buffer in blood and body fluids. Strong acid/conjugate-base pairs like HCl/Cl– do not function as buffers effectively in water.

58 In the Henderson–Hasselbalch relationship, if the ratio [A–]/[HA] increases by a factor of 10, the pH changes by approximately:

A −1 unit

B 0 units

C **+1 unit** ✓

D +10 units

E It becomes exactly 7

▶ **Explanation:** pH = pKa + log([A−]/[HA]). Increasing the ratio by 10 increases the log term by log(10)=1, so pH increases by about 1 unit.

59 **Which statement best describes the common ion effect for HF in water when NaF is added?**

- A HF dissociates more because F− reacts with water to form H^+
- B **HF dissociates less because added F− shifts HF H^+ + F− to the left ✓**
- C HF becomes a strong acid because sodium is a metal
- D The pH must drop because NaF is a salt
- E Nothing changes because equilibrium constants depend only on temperature

▶ **Explanation:** Adding a common ion (F−) increases the product side concentration, so equilibrium shifts left, reducing HF ionization and lowering $[H^+]$. Ka stays the same, but the equilibrium position changes.

60 **Which statement correctly compares conjugate base strengths?**

- A The conjugate base of a strong acid is strong
- B **The conjugate base of a strong acid is weak ✓**
- C All conjugate bases are equally basic
- D A stronger acid always has a stronger conjugate base
- E Conjugate base strength depends only on charge, not on the acid

▶ **Explanation:** Strong acids dissociate almost completely, meaning their conjugate bases have very little tendency to accept protons (they are very weak bases).

61 **Which species is the strongest Brønsted base in water (qualitatively)?**

A Cl^-

B CH_3COO^-

C NH_3

D **OH^-** ✓

E H_2O

▶ **Explanation:** OH^- is a strong Brønsted base in water. Acetate and ammonia are weaker bases; chloride is the conjugate base of a strong acid (very weak base); water is amphiprotic but much weaker as a base than OH^-.

62 **Which statement about the pH of an extremely dilute strong acid solution (e.g., 1.0×10^{-8} M HCl) is most accurate?**

A pH = 8 exactly, because pH = $-\log(1.0\times10^{-8})$

B **pH is close to 7, because water's autoionization contributes significantly to [H_3O^+]** ✓

C pH must be 14 because it is dilute

D pH becomes undefined below 1.0×10^{-7} M

E pH depends only on the acid's molar mass at such low concentration

▶ **Explanation:** At very low acid concentrations near 10^{-7} M, water's own [H_3O^+] (~10^{-7} M at 25°C) is comparable and cannot be ignored. So pH will be slightly below 7, not 8.

63 **Which statement best describes a Lewis base?**

A Electron-pair acceptor

B **Electron-pair donor** ✓

C Proton donor

D Substance with pH < 7 only

E Substance that forms salts only

▶ **Explanation:** A Lewis base donates an electron pair (e.g., NH3). Lewis acids accept electron pairs. Brønsted acids donate protons.

64 **In the reaction BF3 + NH3 → F3B←NH3, which is the Lewis acid and why?**

A NH3, because it donates a proton

B NH3, because it accepts an electron pair

C **BF3, because it accepts an electron pair** ✓

D BF3, because it donates an electron pair

E Neither; this is not an acid–base reaction

▶ **Explanation:** BF3 is electron-deficient (boron has an incomplete octet) and accepts a lone pair from NH3, so BF3 is the Lewis acid and NH3 is the Lewis base.

65 **Which salt solution's pH is mainly determined by whether Ka of the cation (as an acid) is larger or smaller than Kb of the anion (as a base)?**

A NaCl (strong acid + strong base)

B CH3COONa (weak acid + strong base)

C NH4Cl (strong acid + weak base)

D **NH4CH3COO (weak acid + weak base)** ✓

E HCl(aq) (not a salt)

▶ **Explanation:** For salts made from a weak acid and a weak base, both ions hydrolyze, and the pH depends on the relative strengths (Ka vs Kb). For strong/strong, pH ~7; for weak/strong or strong/weak, the direction is determined mainly by the hydrolyzing ion.

66 Which of the following solutions is expected to have pH > 7 (basic) at 25°C?

- A **NaF(aq)** ✓
- B NaCl(aq)
- C NH4Cl(aq)
- D HCl(aq)
- E HNO3(aq)

▶ **Explanation:** F— is the conjugate base of weak acid HF, so it hydrolyzes to produce OH—, making NaF basic. NaCl is neutral; NH4Cl is acidic; HCl and HNO3 are acids.

67 Which reaction best shows a basic oxide reacting with an acid?

- A CO2 + H2O → H2CO3
- B **CaO + 2HCl → CaCl2 + H2O** ✓
- C SO3 + H2O → H2SO4
- D HCl + H2O → H3O$^+$ + Cl—
- E NH3 + H2O → NH4$^+$ + OH—

▶ **Explanation:** CaO is a basic oxide; it reacts with acids to form a salt and water (neutralization-like behavior). CO2 and SO3 are acidic oxides (form acids with water).

68 **Which statement best explains why carbonates ($CO3^{2}-$) react with acids to produce CO2 gas?**

A $CO3^{2}-$ is a strong acid, so it decomposes into CO2

B **Acid protonates carbonate to form carbonic acid (H2CO3), which decomposes to CO2 and H2O** ✓

C CO2 is produced because the acid oxidizes carbon

D CO2 forms because carbonates contain oxygen, which releases CO2

E CO2 is produced only if the acid is strong; weak acids cannot produce CO2

▶ **Explanation:** Carbonate is protonated stepwise to H2CO3, and carbonic acid decomposes to CO2 + H2O. This is why carbonates fizz with acids (even weak acids can do this if enough is present).

69 **In the reaction HCO3− + H2O H2CO3 + OH−, HCO3− is acting as:**

A A Brønsted acid (proton donor)

B **A Brønsted base (proton acceptor)** ✓

C Neither acid nor base

D A Lewis acid only

E A spectator ion

▶ **Explanation:** HCO3− becomes H2CO3 by accepting a proton from water, so it acts as a Brønsted base. Water donates a proton (acts as the acid), producing OH−.

70 **Which statement about the relationship between pH and acidity is correct for a solution at 25°C?**

A A solution with pH 6 is 10 times more acidic than a solution with pH 5

- B **A solution with pH 5 has 10 times higher [H3O$^+$] than a solution with pH 6** ✓
- C pH measures [OH−] directly, not [H3O$^+$]
- D pH differences only matter for strong acids, not weak acids
- E If pH decreases by 2, [H3O$^+$] decreases by a factor of 100

▶ **Explanation:** Each pH unit corresponds to a 10× change in [H3O$^+$]. So going from pH 6 to pH 5 increases [H3O$^+$] by 10× (more acidic).