

AP Chemistry Unit 7 Review

Acids and Bases

Honors Lecture Notes: Unit 7

Definitions

An Arrhenius acid is a H^+ donor, and an Arrhenius base is an OH^- donor. These definitions are not always accurate, though, though, and so instead we use the Bronsted-Lowry definitions. A Bronsted-Lowry acid is the same as an Arrhenius acid- a proton donor. However, a Bronsted-Lowry base is a proton acceptor. By accepting protons from water, hydroxide ions are created, causing the solution to be basic.

Substances that can act as either acids or bases (such as water) are called amphoteric. Whether an amphoteric substance acts as an acid or as a base depends on what it is combined with.

A conjugate pair is two species that differ from each other by a single proton. Examples: NH_3/NH_4^+ , $HClO_2/ClO_2^-$. The stronger the acid or base is, the weaker the conjugate is.

Ions In Solution

Typically, most anions create basic solutions, as due to their negative charge they will accept protons from water readily. The exception to this rule is the conjugates of the strong acids. As strong acids dissociate completely, their conjugate ions tend to be ineffective bases. The strong acids are: HCl , HNO_3 , H_2SO_4 , HI , HBr , and $HClO_4$.

Ex: Na_2CO_3 is basic in solution because:



$NaCl$, however, is neutral in solution, as Cl^- is the conjugate of HCl and thus will not accept protons from water.

Unlike anions, cations are generally not considered to be acidic as they don't often have protons to donate (notable exceptions: NH_4^+). Strong bases are defined as those which donate 100% of their OH^- ions in solution. Strong bases are any alkali hydroxide as well as $Ba(OH)_2$ and $Sr(OH)_2$.

pH/pOH

To determine the relative acidity/basicity of a solution, the pH/pOH is used. $pH = -\log[H^+]$ and $pOH = -\log[OH^-]$. At $25^\circ C$, $pH + pOH = 14$ and $[H^+][OH^-] = 1.0 \times 10^{-14}$ for any solution. Thus, as $[H^+]$ goes up, pH goes down. Likewise, as pH goes down, pOH goes up.

Any solution with a pH of less than 7 is considered acidic, and a pH of above 7 indicates a basic solution. Note that pH/pOH are NOT limited to a 0-14 scale. Sufficiently concentrated strong acids can have negative pHs, and concentrated strong bases can have pHs in excess of 14.

Auto-ionization of Water



For water at $25^\circ C$, $K_w = 1.0 \times 10^{-14}$, and $[H^+] = [OH^-] = 1.0 \times 10^{-7} M$. The K_w of water (like any equilibrium constant) changes with temperature. The auto-ionization of water is an endothermic process, and as such, if temperature increases K_w increases. This means that at temperatures above $25^\circ C$, the pH of pure water is NOT 7.0. However, neither is the pOH. For instance, at $50^\circ C$, $K_w = 5.5 \times 10^{-14}$, and thus $[H^+] = [OH^-] = 2.3 \times 10^{-7} M$. At $50^\circ C$, $pH = pOH = 6.63$, and $pH + pOH = 13.2$.

K_a/K_b

The strength of a weak acid or base is defined by the equilibrium constant. These equilibrium constants are calculated the same way as any other equilibrium constant [products/reactants], but as it is always the loss/gain of a single proton, there are never any exponents. K_a/K_b values are often used to calculate the pH of a solution containing a weak acid or base.

Ex: Calculate the pH of 1.5 M acetic acid, $HC_2H_3O_2$. ($K_a = 1.8 \times 10^{-5}$).



$$K_a = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

As H^+ and $C_2H_3O_2^-$ are in a 1:1 ratio, both values can be replaced by x . So: $1.8 \times 10^{-5} = \frac{x^2}{1.5}$

$$x = [H^+] = 0.0051 M \quad -\log(0.0051 M) = 2.28$$

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Table of Contents Chemistry Unit 7 Review Sheet Answer Key

1. Understanding the eBook Chemistry Unit 7 Review Sheet Answer Key
 - The Rise of Digital Reading Chemistry Unit 7 Review Sheet Answer Key
 - Advantages of eBooks Over Traditional Books
2. Identifying Chemistry Unit 7 Review Sheet Answer Key
 - Exploring Different Genres
 - Considering Fiction vs. Non-Fiction
 - Determining Your Reading Goals
3. Choosing the Right eBook Platform
 - Popular eBook Platforms
 - Features to Look for in an Chemistry Unit 7 Review Sheet Answer Key
 - User-Friendly Interface
4. Exploring eBook Recommendations from Chemistry Unit 7 Review Sheet Answer Key
 - Personalized Recommendations
 - Chemistry Unit 7 Review Sheet Answer Key User Reviews and Ratings
 - Chemistry Unit 7 Review Sheet Answer Key and Bestseller Lists
5. Accessing Chemistry Unit 7 Review Sheet Answer Key Free and Paid eBooks
 - Chemistry Unit 7 Review Sheet Answer Key Public Domain eBooks
 - Chemistry Unit 7 Review Sheet Answer Key eBook Subscription Services
 - Chemistry Unit 7 Review Sheet Answer Key Budget-Friendly Options
6. Navigating Chemistry Unit 7 Review Sheet Answer Key eBook Formats

- ePub, PDF, MOBI, and More
 - Chemistry Unit 7 Review Sheet Answer Key Compatibility with Devices
 - Chemistry Unit 7 Review Sheet Answer Key Enhanced eBook Features
7. Enhancing Your Reading Experience
- Adjustable Fonts and Text Sizes of Chemistry Unit 7 Review Sheet Answer Key
 - Highlighting and Note-Taking Chemistry Unit 7 Review Sheet Answer Key
 - Interactive Elements Chemistry Unit 7 Review Sheet Answer Key
8. Staying Engaged with Chemistry Unit 7 Review Sheet Answer Key
- Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Chemistry Unit 7 Review Sheet Answer Key
9. Balancing eBooks and Physical Books Chemistry Unit 7 Review Sheet Answer Key
- Benefits of a Digital Library
 - Creating a Diverse Reading Collection Chemistry Unit 7 Review Sheet Answer Key
10. Overcoming Reading Challenges
- Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
11. Cultivating a Reading Routine Chemistry Unit 7 Review Sheet Answer Key
- Setting Reading Goals Chemistry Unit 7 Review Sheet Answer Key
 - Carving Out Dedicated Reading Time
12. Sourcing Reliable Information of Chemistry Unit 7 Review Sheet Answer Key
- Fact-Checking eBook Content of Chemistry Unit 7 Review Sheet Answer Key
 - Distinguishing Credible Sources
13. Promoting Lifelong Learning
- Utilizing eBooks for Skill Development
 - Exploring Educational eBooks
14. Embracing eBook Trends
- Integration of Multimedia Elements
 - Interactive and Gamified eBooks

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