

Calculations In Chemistry

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Module 29: Acid-Base Fundamentals

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Module 29 — Acid-Base Fundamentals

Timing: The naming of acids was covered in Lesson 7D. Acid-base *neutralization* calculations were covered in Module 14. Begin this module when you are assigned calculations that include K_w or pH.

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Lesson 29A: Acid-Base Math Review

Pretest: This lesson reviews math that is used in acid-base calculations. If your exponential math is good, try the *last* lettered part of the **Practice** questions 1-4. If you get those right, you may skip this lesson.

If more review is needed, study the rules below, then do more of the practice set. If more detailed review is needed, see Lessons 1A to 1C.

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Rules For Exponential Notation

- To *multiply* exponential terms, *add* the exponents.
- To *divide* exponentials, *subtract* the exponents.
- Numbers expressed in exponential notation have three parts:

-	5.25	x	10 ³
	^		^
	significand		exponential
- Answers in exponential notation should be converted to scientific notation. This places the decimal in the significand after the first digit that is not a zero.
 - Moving the decimal point Y times to make the significand *larger*, make the exponent of 10 *smaller* by a count of Y.
 - Moving the decimal point Y times to make the significand *smaller*, make the exponent *larger* by a count of Y.

“If you make one larger, make the other smaller.”
- When changing fixed decimal numbers to exponential notation, the number of times you move the decimal becomes the positive or negative number of the exponential.
- In scientific notation, numeric values larger than one will have positive powers of 10. Values between zero and one (0.XX) will have negative powers of 10.
- In calculations using exponential notation, handle the two parts separately: do number math by number rules and exponential math by exponential rules.
- If an exponential term does not have a number in front, add a “1 x” in front of the exponential so that the number-number division is clear.

Practice

Do the problems below without looking at the rules. If you find that you need to look back, write a summary of the rules above, recite your rules, *then* continue with the practice.

To speed your progress, do every second or third problem. If you cannot solve easily, do more parts of that problem. Check your answers at the end of this lesson.

Try problems 1-3 without a calculator.

1. If you need help, see Rule 1.
 - a. $(10^{-8})(10^{+2}) =$
 - b. $(10^{-3})(10^{-12}) =$
 - c. $(x)(10^{-12}) = 10^{-14}$; $x =$
 - d. $[H^+](10^{-9}) = 10^{-14}$; $[H^+] =$
 - e. $(10^{-3}) [OH^-] = 10^{-14}$; $[OH^-] =$

2. *Try* to solve these without a calculator. Convert your final answer to scientific notation. Need a hint? See Rule 7. Take the paper you need for careful work: solve in your notebook if needed.
 - a. $(2.0 \times 10^1)(3.0 \times 10^{-11}) =$
 - b. $\frac{1.0 \times 10^{-14}}{2.0 \times 10^4} =$
 - c. $\frac{1.0 \times 10^{-14}}{3.0 \times 10^{-4}} =$
 - d. $(x)(2.0 \times 10^{-8}) = 10. \times 10^{-15}$; $x =$
 - e. $(2.5 \times 10^{-2})(x) = 10. \times 10^{-15}$; $x =$

3. *Try* to solve without a calculator. It may help to convert all values to scientific notation. Assume a value that is written as only an exponential is exact.
 - a. $\frac{1.0 \times 10^{-14}}{0.040} =$
 - b. $\frac{1.0 \times 10^{-14}}{0.0030} =$
 - c. $(x)(0.20) = 1.0 \times 10^{-14}$; $x =$
 - d. $(0.0125)(x) = 10^{-14}$; $x =$

4. Use a calculator for the numbers but not the exponentials. Convert answers to scientific notation.

a. $\frac{1.0 \times 10^{-14}}{3.25 \times 10^{-5}} =$

b. $\frac{1.0 \times 10^{-14}}{8.8 \times 10^{-4}} =$

c. $\frac{10^{-14}}{2.4} =$

d. $\frac{10^{-14}}{4.3 \times 10^{-4}} =$

e. $(x)(6.7 \times 10^{-12}) = 1.0 \times 10^{-14}; x =$

f. $(1.25 \times 10^{-7})(x) = 10^{-14}; x =$

ANSWERS

1a. $(10^{-8})(10^{+2}) = 10^{-6}$ 1b. 10^{-15} 1c. 10^{-2} 1d. 10^{-5}

1e. $(10^{-3})[\text{OH}^{-}] = 10^{-14}; [\text{OH}^{-}] = 10^{-11}$

2a. 6.0×10^{-10} 2b. 5.0×10^{-19} 2c. 3.3×10^{-11}

2d. 5.0×10^{-7} 2e. 4.0×10^{-13}

3a. 2.5×10^{-13} 3b. 3.3×10^{-12} 3c. 5.0×10^{-14}

3d. $x = \frac{1.0 \times 10^{-14}}{1.25 \times 10^{-2}} = 0.80 \times 10^{-12} = 8.0 \times 10^{-13}$

4a. $\frac{1.0 \times 10^{-14}}{3.25 \times 10^{-5}} = 0.31 \times 10^{-9} = 3.1 \times 10^{-10}$

4b. $\frac{1.0 \times 10^{-14}}{8.8 \times 10^{-4}} = 0.11 \times 10^{-10} = 1.1 \times 10^{-11}$

4c. $\frac{10^{-14}}{2.4} = \frac{1.0 \times 10^{-14}}{2.4} = 4.2 \times 10^{-15}$ 4d. 2.3×10^{-11} 4e. 1.5×10^{-3}

4f. $x = \frac{1.0 \times 10^{-14}}{1.25 \times 10^{-7}} = 0.80 \times 10^{-7} = 8.0 \times 10^{-8}$

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Lesson 29B: K_w Calculations: H^+ and OH^-

Acid-Base Review

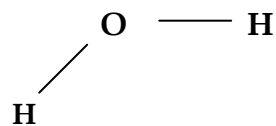
In Module 14, we considered acid-base neutralization calculations. In those problems, the question was: if we have a known amount of an acid or base, can we find its stoichiometric equivalent: the amount of the opposite that is needed to exactly neutralize the acid or base? These “reaction amount” calculations for stoichiometric equivalents were solved using conversion stoichiometry.

Beginning in this module, we return to acids and bases to ask additional questions that are important in both chemistry and biology. What is the nature of acid and base solutions before they react? Which particles and ions are present? How are concentrated acid or base solutions different from dilute solutions, and how do strong acids like hydrochloric acid differ from weak acids such as vinegar that we frequently consume as food?

In most cases, we are interested in the behavior of acids and bases when they are dissolved in an aqueous solution. Let's start with the molecule that is present in the highest concentration in aqueous solutions: H_2O .

The Ionization of Water

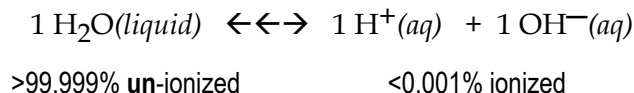
A water molecule oxygen in the middle, two bonds, and a bent shape.



At temperatures above absolute zero, the bonds in water bend and stretch. At room temperature, the liquid molecules also move and collide at high average speeds. In part as a result of these motions, for about one in 500 million molecules at room temperature, one of the bonds in liquid water is broken. The result is the formation of two ions, H^+ and OH^- , that also move about in the water solution.

When a compound divides into ions that can move about freely, the reaction is termed **dissociation**. When the initial compound has covalent character, the process can also be termed **ionization** (forming ions). In practice, all compounds have some covalent character, and the terms dissociation and ionization are often used interchangeably.

This separation of water into ions is reversible and can be represented by the equation



In pure (**distilled**) liquid water, the *number* of H^+ and OH^- ions must be equal, since they must be formed in a 1 to 1 ratio when water ionizes. The *concentration* of the H^+ and OH^- ions must also be equal, since the equal moles of the two ions will be contained in the same volume of solution. At room temperature, these *ion* concentrations are very small: 1.0×10^{-7} moles per liter for each. The concentration of the un-ionized water molecules is very large in comparison: about 55 mol/L.

However, even at these low concentrations, small changes in the balance between H^+ and OH^- can have a large influence on reactions that are important in chemistry and biology.

Water's Ionization Constant: K_w

For the reversible ionization of water: $\text{H}_2\text{O}(\text{liquid}) \leftrightarrow \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq})$

the equilibrium constant expression *could* be written as:

$$K = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}$$

However, in aqueous solutions (those where water is the solvent), the concentration of the *non*-ionized water molecules is high (about 55 M) for the relatively dilute solutions used in most lab experiments, and it does not change substantially during dilution or reactions.

In equilibrium constants, concentrations that remain close to constant during reactions are generally included as part of the numeric value for K and represented by a 1 in the K expression. The result is a simplified equilibrium constant expression (labeled K_w) for the relationship between H^+ and OH^- in an aqueous solution:

$$K_w = [\text{H}^+][\text{OH}^-]$$

At room temperature (25°C), the value for K_w is 1.0×10^{-14} . This small numeric value (0.000 000 000 000 010) indicates that the reaction favors the reactants: very few water molecules separate into ions.

As temperature increases, the molecules of water collide with higher energy, the bonds bend and stretch more vigorously, and the bonds break more often. Of importance in biology, at *body temperature* in mammals (37°C), the $[\text{H}^+]$ in water is about 1.6×10^{-7} as opposed to 1.0×10^{-7} at 25°C.

However, for acid-base calculations in *chemistry*, you may assume that conditions are at 25°C unless otherwise noted. In acid-base calculations, we will use this K_w equation:

$$K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14} \quad \text{at } 25^\circ\text{C}$$

This relationship between $[\text{H}^+]$ and $[\text{OH}^-]$ is an inverse proportion. If substances are added to water to make one ion concentration increase, the other must decrease in the same proportion: if one triples, the other must become 1/3 of its original concentration.

Acid-Base Terminology

In pure water, the concentration of H^+ and OH^- ions must be equal. However, if acids or bases are dissolved into water, this balance is upset.

By what are termed the **classical** definitions of acids and bases:

- An **acid** is a substance that ionizes in water and forms $[\text{H}^+]$ ions.
- A **base** is a substance that ionizes in water and forms $[\text{OH}^-]$ ions.

These are also called the **Arrhenius** definitions for acids and bases, after the Swedish chemist Svante Arrhenius who first proposed the existence of electrically charged particles (ions).

When acids or bases are added to water, the water continues to ionize slightly, and the K_w equation will continue to predict the relationship between the $[H^+]$ and $[OH^-]$ ions. This means an acidic or basic solution, if either $[H^+]$ or $[OH^-]$ is known, the concentration of the other ion can be calculated using the K_w equation.

A useful rule is: in calculations that include *both* $[H^+]$ and $[OH^-]$ in the WANTED and DATA, write the K_w equation. We will call this the

$$\underline{K_w \text{ prompt:}} \quad \text{See } [H^+] \text{ and } [OH^-]? \text{ Write: } K_w = \boxed{[H^+][OH^-] = 1.0 \times 10^{-14}}$$

Some problems will ask for the approximate $[H^+]$ and $[OH^-]$ in acid or base solutions, and those calculations can be done by mental arithmetic. Try this problem.

Q. In a solution with a $[OH^-]$ of about 10^{-2} M, what is the $[H^+]$?

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Answer

WANT: $[H^+]$

DATA: $[OH^-] \approx 10^{-2}$ M

Prompt: See $[H^+]$ and $[OH^-]$? Write: $K_w = \boxed{[H^+][OH^-] = 1.0 \times 10^{-14}}$

Substitute the known $[OH^-]$:

$$[H^+](\sim 10^{-2}) = 10^{-14} ; \text{ by inspection, } [H^+] \text{ must } \boxed{\approx 10^{-12} \text{ M}}$$

Note that as always when solving K equations, *units* are omitted during calculations, but if the WANTED unit is a concentration, the unit mol/L (**M**) is added to the answer.

In other problems, you will need to calculate the ion concentrations more precisely. Try this calculation in your notebook, then check your answer below.

Q. If the $[H^+]$ in an aqueous solution is 5.0×10^{-3} M, find the $[OH^-]$.

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WANT: $[OH^-] = ?$

DATA: $[H^+] = 5.0 \times 10^{-3}$ M

$$K_w = \boxed{[H^+][OH^-] = 1.0 \times 10^{-14}} \quad (K_w \text{ prompt})$$

SOLVE: Solve the equation for the wanted *symbol*, then plug in the DATA.

$$? = [OH^-] = \frac{1.0 \times 10^{-14}}{[H^+]} = \frac{1.0 \times 10^{-14}}{5.0 \times 10^{-3}} = 0.20 \times 10^{-11} = \boxed{2.0 \times 10^{-12} \text{ M}}$$

K_w Check

When using the K_w equation, do a *check* at the end: *estimate* $[H^+]$ times $[OH^-]$ (circled above). The result must equal 10.0×10^{-15} or 1.0×10^{-14} . Try multiplying the circled values above in your head. Does the answer check?

Practice: Do parts b and d of both, plus problem 3. Need more practice? Do more parts.

- In these aqueous solutions, find the $[H^+]$ if the $[OH^-]$ is
 - 10^{-11} moles per liter
 - 0.010 Molar
 - 5.0×10^{-11} M
 - 0.036 M
- In these aqueous solutions, find the $[OH^-]$ if the $[H^+]$ is
 - 10^{-9} moles per liter
 - 1.0 Molar
 - 3.0×10^{-5} M
 - 1.25 M
- If 20.0 millimoles of OH^- ions are dissolved in 400. mL of solution, find
 - $[OH^-]$
 - $[H^+]$.

ANSWERS

- Since $[H^+][OH^-] = 1.0 \times 10^{-14}$; $[H^+](10^{-11}) = 10^{-14}$; $[H^+] = 10^{-3}$ M
 - 0.010 Molar = $[OH^-]$; $[H^+] = 1.0 \times 10^{-2}$ M
 $[H^+][OH^-] = 1.0 \times 10^{-14}$; $[H^+](1.0 \times 10^{-2}) = 1.0 \times 10^{-14}$; $[H^+] = 1.0 \times 10^{-12}$ M
 - 5.0×10^{-11} M = $[OH^-]$; WANT $[H^+]$; the equation relating the two symbols is
 $[H^+][OH^-] = 1.0 \times 10^{-14}$. Solving for the wanted symbol:

$$[H^+] = \frac{1.0 \times 10^{-14}}{[OH^-]} = \frac{1.0 \times 10^{-14}}{5.0 \times 10^{-11}} = 0.20 \times 10^{-3} = 2.0 \times 10^{-4} \text{ M } H^+$$
 - 0.036 M = $[OH^-]$; WANT $[H^+]$; the equation using those symbols is $[H^+][OH^-] = 1.0 \times 10^{-14}$

$$[H^+] = \frac{1.0 \times 10^{-14}}{[OH^-]} = \frac{1.0 \times 10^{-14}}{3.6 \times 10^{-2}} = 0.28 \times 10^{-12} = 2.8 \times 10^{-13} \text{ M } H^+$$
- 10^{-9} moles per liter = $[H^+]$, want $[OH^-]$, the equation relating those symbols is
 $[H^+][OH^-] = 1.0 \times 10^{-14}$; $(10^{-9}) [OH^-] = 10^{-14}$; $[OH^-] = 10^{-5}$ M
 - 1.0 Molar = $[H^+]$, want $[OH^-]$, that calls the K_w prompt:
 write $[H^+][OH^-] = 1.0 \times 10^{-14}$; $(1.0) [OH^-] = 10^{-14}$; $[OH^-] = 1.0 \times 10^{-14}$ M
 - 3.0×10^{-5} M = $[H^+]$, want $[OH^-]$, write $[H^+][OH^-] = 1.0 \times 10^{-14}$;

$$[OH^-] = \frac{1.0 \times 10^{-14}}{[H^+]} = \frac{1.0 \times 10^{-14}}{3.0 \times 10^{-5}} = 0.33 \times 10^{-9} = 3.3 \times 10^{-10} \text{ M } OH^-$$

$$d. \quad 1.25 \text{ M} = [\text{H}^+]; \quad [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{[\text{H}^+]} = \frac{1.0 \times 10^{-14}}{1.25} = 0.80 \times 10^{-14} = \boxed{8.0 \times 10^{-15} \text{ M OH}^-}$$

$$3a. \quad \text{WANTED} = ? = [\text{OH}^-] = \frac{\text{mol OH}^-}{\text{L soln.}}$$

DATA: 20.0 mmol KOH = 400. mL soln. (two measures of same solution)
(If you WANT a ratio, all of the DATA will be in equalities.)

$$\text{SOLVE: } ? \frac{\text{mol OH}^-}{\text{L soln.}} = \frac{20.0 \text{ mmol OH}^-}{400. \text{ mL soln.}} = \mathbf{0.0500 \frac{\text{mol OH}^-}{\text{L soln.}}}$$

(Since a prefix is an abbreviation for an exponential, the same prefix on the top and bottom can cancel.)

3b. Knowing $[\text{OH}^-]$, $[\text{H}^+]$ can be solved using the K_w prompt.)

$$\text{Write } \boxed{[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}}$$

$$[\text{H}^+] = \frac{1.0 \times 10^{-14}}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{5.00 \times 10^{-2}} = 0.20 \times 10^{-12} = \boxed{2.0 \times 10^{-13} \text{ M H}^+}$$

* * * * *

Lesson 29C: Strong Acid Solutions

Prerequisites: If you have difficulty with the calculations in this lesson, review Lesson 12B.

* * * * *

Definitions

By the classical (Arrhenius) definitions in chemistry, a compound that is a

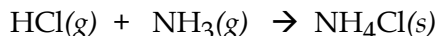
- **Strong acid** ionizes essentially 100% in water and forms an H^+ ion;
- **Strong base** ionizes essentially 100% in water and forms an OH^- ion;
- **Weak acid or base** dissociates (ionizes) only partially when dissolved in water.

An H^+ ion has one proton and no electrons, so the H^+ ion is often referred to as a *proton*. Acids can be classified as **monoprotic** (containing *one* hydrogen atom that can ionize) or **polyprotic** (containing more than one).

Acid-Base States

Molecules that are acids and bases can react in a variety of ways.

For example, one type of acid-base reaction is the mixing of hydrogen chloride gas and ammonia gas to form solid ammonium chloride.



However, *most* acid-base reactions are conducted in water, and in these lessons, if no state for an acidic or basic particle is shown, assume that the state is aqueous (*aq*).

Strong Acids

In chemistry, we often have a need for strong acid solutions. The strong acids most frequently encountered are HCl, HNO₃, and H₂SO₄.

- HCl and HNO₃ are strong *monoprotic* acids. Both are highly soluble in water and dissociate essentially 100% to release one H⁺ ion.

Other strong monoprotic acids include HBr, HI, HClO₄, and HMnO₄, but because these substances may undergo redox as well as acid-base reactions, they are used less often for reactions that require a strong acid.

- H₂SO₄ is a strong *diprotic* acid. When H₂SO₄ is dissolved in water, the first proton ionizes essentially 100%, but the second ionizes only partially.

The ionization of acids is more complex than the ionization of other ionic compounds in part because the bond of the H in acids has a balance of ionic and covalent character. Because of the mixed nature of bonds to hydrogen, to calculate ion concentrations in acidic solutions, we need three sets of rules:

- one for strong acids (such as HCl and HNO₃) that ionize completely,
- one for weak acids, in which ionization goes to equilibrium, and
- one for polyprotic acids (such as H₂SO₄) in which some H atoms ionize more easily than others.

Our rules for bases will parallel these acid rules. Let us start with the rules for solutions of the strong acids HCl and HNO₃.

Ion Concentrations In Strong Acid Solutions

The concentration of an HCl or HNO₃ solution is usually expressed in terms of an *un-ionized* acid formula, such as [HCl] = 0.25 M. However, the un-ionized formula does not represent the particles that are actually present in the solution.

For example:

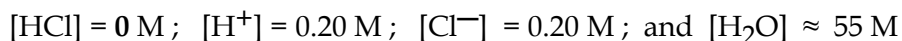
The gas hydrogen chloride (HCl) readily dissolves in water to form a solution of *hydrochloric acid*. If 0.20 moles of HCl is dissolved per one liter of solution, the solution concentration is written as “[HCl] = 0.20 M” based on how much HCl is added when mixing the solution. We will call this 0.20 M the [HCl]**as mixed**.

However, there are *no* particles of HCl present in an “HCl solution.” Why? As HCl dissolves in water, it immediately separates into ions:



This reaction is reversible, but the right side is so strongly favored at equilibrium that in water, essentially all of the HCl is converted to ions.

In a solution *labeled* [HCl] = 0.20 M, the actual concentrations of the particles are



In addition, a very small amount of OH^- ion is present in the solution (more on that later).

Calculating Ion Concentrations In Strong Acid Solutions

The ionization of a strong acid parallels what happens to other ionic compounds that separate into ions essentially 100% when dissolved in water (see Lesson 12B). The key rule for calculations involving such compounds is:

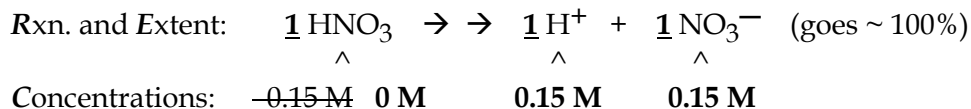
For Substances Separating 100% Into Ions, Write the REC Steps

In calculations involving [ions], if a substance ionizes ~100%,

- **R:** Write the balanced ionization **R**reaction equation. After the equation, write
- **E:** The **E**xtent of the reaction (“goes ~100%”). Below each particle, write
- **C:** The **C**oncentration of the particle, based on the *coefficient* ratios.

If a reaction goes to completion, the coefficients supply the *mole* ratios of reactants used up and products formed. For the 100% dissociation of strong acids, since all of the reaction particles are in the same constant volume of solution, the coefficients also supply the *mole per liter* reaction ratios.

Example: In 0.15 M HNO_3 , what are the [ions]? Write the REC steps.



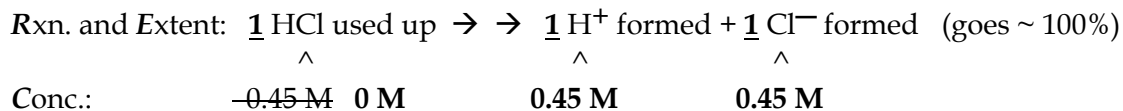
In your notebook apply the REC steps to this problem.

Q. In a 0.45 M HCl solution, write the

- a. $[\text{HCl}]_{\text{as mixed}}$ b. $[\text{H}^+]_{\text{in soln.}}$ c. $[\text{Cl}^-]_{\text{in soln.}}$

* * * * *

Answer: Since HCl is a strong acid, in water it ionizes ~100%. To find ion concentrations for substances that ionize ~100%, write the REC steps.



The bottom row shows the WANTED H^+ and Cl^- concentrations. The $[\text{HCl}]_{\text{as mixed}}$ was 0.45 M, but in solution is 0 M.

In some problems, to find [ions], the $[\text{HCl}]$ or $[\text{HNO}_3]_{\text{as mixed}}$ must be solved first. In your notebook, try:

Q. If 0.030 moles of HNO_3 is mixed with water to form 150 mL of solution, find

- a. $[\text{HNO}_3]_{\text{as mixed}}$ b. $[\text{H}^+]_{\text{in soln.}}$ c. $[\text{NO}_3^-]_{\text{in soln.}}$

* * * * *

When a strong acid is dissolved in water, H^+ is contributed by *both* the acid and the water. However, if the strong acid is mixed in any significant concentration, the share of H^+ ions contributed by the water can be ignored.

Why? Consider a 0.20 M HCl solution. The strong acid contributes 0.20 moles of H^+ ions per liter to the water. Before the acid was added, the water contained only 10^{-7} moles of H^+ ions per liter. If we examine the uncertainty in these two amounts:

0.20 M H^+ from the acid, with doubt in the hundredth's place, compares to
0.0000001 M H^+ initially in the water, with doubt in a much lower place.

This is one indication that any initial H^+ contribution from the water's ions is too small to be significant. The acid ionization is the *dominant* reaction. For this reason, the rule is:

In an acid solution, use acid ionization rules to find $[H^+]$

The HCl and HNO₃ Quick Rule

REC steps show concentrations for all ions formed when substances ionize 100%. However, in many strong acid calculations, only $[H^+]$ is wanted, and you can use this

Quick rule: $[HCl]_{\text{mixed}}$ or $[HNO_3]_{\text{mixed}} = [H^+]_{\text{in solution}}$

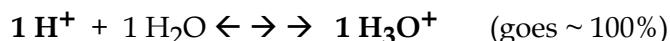
Example: In a solution labeled 0.35 M HCl, $[H^+]_{\text{in soln.}} = 0.35 \text{ M}$

When a *monoprotic* strong acid dissolves in water, the ratio of the [acid as mixed but used up] and the $[H^+]$ that forms in solution is always **1 to 1**.

H⁺ and H₃O⁺ Ions -- Equivalent

An H^+ ion has one proton and no electrons, so it is often referred to as a proton.

In aqueous solutions, the proton released by an acid is nearly always is found attached to a water molecule, forming a **hydronium ion** (H_3O^+). This reaction can be represented as



Textbooks often show acids in water forming H^+ in some reactions and H_3O^+ in others. In calculations involving acids in aqueous solutions, the symbols H^+ and H_3O^+ in most cases are considered to be equivalent. When H_3O^+ is encountered in calculations, use this rule:

If you see H_3O^+ , write: $H_3O^+ = H^+$

Let's list the rules learned so far for acids. Design flashcards as needed that will help you to apply these rules from memory.

Summary: Acid Rules

1. See $[H^+]$ and $[OH^-]$? Write: $K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$
2. Strong monoprotic acids ionize ~100% to form H^+ . For [ions], write the *REC* steps.
3. Quick rule: $[HCl \text{ or } HNO_3]_{\text{mixed}} = [H^+]_{\text{in soln.}}$
4. See H_3O^+ ? Write: $H_3O^+ = H^+$
5. In an acid solution, use acid ionization rules to find $[H^+]$.

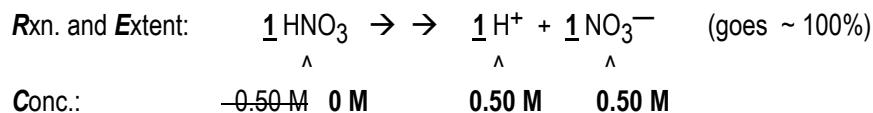
Practice B: Memorize the rules above, *then* do these problems.

1. In 0.25 M HNO_3 a. $[H^+] =$ b. $[H_3O^+] =$
2. In 2.0 M HCl , a. $[H_3O^+] =$ b. $[H^+] =$
3. In an HCl solution, what is the formula for the particle that is the conjugate base?

ANSWERS

Practice A

1a,b. For problems involving HCl or HNO_3 and [ions], write the *REC* steps for 100% ionization.



2a. WANT: $\frac{? \text{ mol HCl}}{\text{L soln}} = [HCl]_{\text{as mixed}}$ (it helps to label units with the symbols used in rules and equations)

DATA: 7.30 g HCl = 250 mL soln (equivalent: two measures; same soln.)

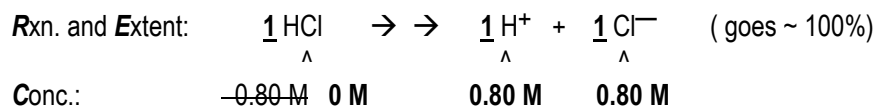
36.5 g HCl = 1 mol HCl (grams prompt)

You want moles over liters. The data includes grams, moles and mL. Use conversions. WANT a ratio? Start with a ratio.

$$\text{SOLVE: } \frac{? \text{ mol HCl}}{\text{L soln}} = \frac{7.30 \text{ g HCl}}{250 \text{ mL soln}} \cdot \frac{1 \text{ mol HCl}}{36.5 \text{ g HCl}} \cdot \frac{1 \text{ mL}}{10^{-3} \text{ L}} = \boxed{0.80 \text{ M HCl} = [HCl]_{\text{mixed}}}$$

When solving in *parts*, label answers with *symbols* from rules that may be needed in later parts.

2b,c. For problems involving HCl or HNO_3 and [ions], solve by the *REC* steps.



$$[\text{HCl}]_{\text{mixed}} = \boxed{0.80 \text{ M from part a} = [\text{H}^+]_{\text{in soln}} = [\text{Cl}^-]_{\text{in soln}}}$$

Practice B

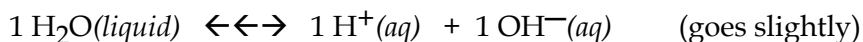
Quick rule: In HCl or HNO₃, for [H⁺], use $\boxed{[\text{HCl or HNO}_3]_{\text{mixed}} = [\text{H}^+]_{\text{in soln.}}}$

- In 0.25 M HNO₃ a. [H⁺] = 0.25 M b. [H₃O⁺] = [H⁺] = 0.25 M
- In 2.0 M HCl, a. [H₃O⁺] = [H⁺] = 2.0 M b. [H⁺] = 2.0 M
- Cl⁻. The conjugate base has one fewer H atoms and one fewer positive charges than the acid from which it is formed.

* * * * *

Lesson 29D: The [OH⁻] in Strong Acid Solutions**The Impact of a Strong Acid on the [OH⁻] in Water**

The ionization of an acid in water affects the other reaction occurring in an acid solution: the reversible “auto-ionization” of water.



Adding acid to water shifts this equilibrium in accord with Le Chatelier’s Principle. As the [H⁺] in the solution increases, the equilibrium shifts toward the side of the reaction equation opposite the H⁺, decreasing the [OH⁻]. The *value* of the [OH⁻] after the shift can be calculated by the equilibrium constant for water’s ionization:

$$K_{\text{W}} = \boxed{[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}} \text{ at } 25^\circ\text{C}$$

When solving for [H⁺] and [OH⁻] in an acid solution, let us add to Rule 5 as follows.

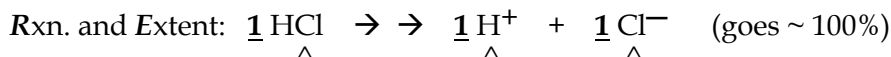
5. In an acid solution, use acid ionization rules to find [H⁺], then K_{W} to find [OH⁻].

Apply the rule to this problem.

Q. In a 0.40 M HCl solution, find a. [H⁺] b. [Cl⁻] c. [OH⁻]

* * * * *

- a.b. You may be able to solve parts a and b by the logic of the quick steps, or you can solve methodically using the rule: to find [ions] in a strong acid solution, write the REC steps.



- c. WANT [OH⁻].

$\boxed{\text{In an acid solution, use acid rules to find [H}^+], \text{ then } K_{\text{W}} \text{ to find [OH}^-].}$

Part a found [H⁺] = 0.40 M To find [OH⁻], use K_{W} .

$$K_w = \boxed{[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}};$$

SOLVE the boxed equation for the symbol WANTED.

$$? = [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{[\text{H}^+]} = \frac{1.0 \times 10^{-14}}{0.40} = \boxed{2.5 \times 10^{-14} \text{ M}}$$

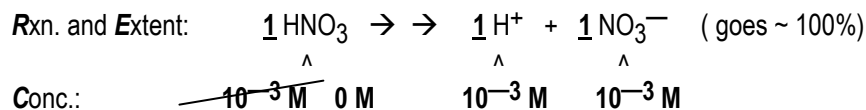
K_w check: $[\text{H}^+] \times [\text{OH}^-]$ must $\approx 10.0 \times 10^{-15}$ or 1.0×10^{-14} . Check.

Practice

- From memory, write the 5 numbered acid-base rules learned so far in this module.
- In a solution labeled 10^{-3} M HNO_3 , find the
 - $[\text{H}^+]$
 - $[\text{NO}_3^-]$
 - $[\text{H}_3\text{O}^+]$
 - $[\text{OH}^-]$
- Solve in your notebook: In a solution labeled 0.020 M HCl ,
 - Write the balanced equation for the ionization of this strong acid.
 - Which side will be favored in this ionization: products, or reactants?
 - What is the *mole to mole* ratio between HCl used up and H^+ formed?
 - What is the *mol/L* ratio between HCl used up and H^+ formed?
 - $[\text{H}^+] = ?$
 - $[\text{Cl}^-] = ?$
 - $[\text{OH}^-] = ?$
 - $[\text{H}_3\text{O}^+] = ?$

ANSWERS

2 a,b. For problems involving HCl or HNO_3 and [ions], solve by the *REC* steps.



2c. $10^{-3} \text{ M} = [\text{H}^+] = [\text{H}_3\text{O}^+]$

2d. To find $[\text{OH}^-]$ in an acid solution, use acid rules to find $[\text{H}^+]$, then K_w to find $[\text{OH}^-]$.

Since $[\text{H}^+] = 10^{-3} \text{ M}$ and $[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$; $[\text{OH}^-] = 10^{-11} \text{ M}$.

3a,b. Equation: $1 \text{ HCl}(g) \rightarrow \rightarrow 1 \text{ H}^+(aq) + 1 \text{ Cl}^-(aq)$ (goes 100%) **Products favored**

3c. 1 mol HCl used up = 1 mole H^+ formed.

3d. Since all moles are found in the same liters of solution, the mole and mol/L ratios are the same: **1 to 1**.

1 mol/L HCl used up = 1 mol/L H^+ formed, which can be written: $[\text{HCl}]_{\text{used up}} = [\text{H}^+]_{\text{formed}}$

3e, f. $[\text{HCl}]_{\text{as mixed}} = 0.020 \text{ M} = [\text{H}^+]_{\text{in soln.}} = [\text{Cl}^-]_{\text{in soln.}}$

3g. $[\text{OH}^-]$? In an acid solution, use acid rules to find $[\text{H}^+]$ (done above), then use K_W to find $[\text{OH}^-]$.

Since $[\text{H}^+] = 0.020 \text{ M} = 2.0 \times 10^{-2} \text{ M}$, and $[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$;

$$[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{[\text{H}^+]} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-2}} = 0.50 \times 10^{-12} = \mathbf{5.0 \times 10^{-13} \text{ M}}$$

K_W check: $[\text{H}^+][\text{OH}^-]$ (circled above) must estimate to 10.0×10^{-15} or 1.0×10^{-14}

$[\text{H}^+][\text{OH}^-] = 2 \times 5 = 10$; $10^{-2} \times 10^{-13} = 10^{-15}$, combined = 10×10^{-15} . Check.

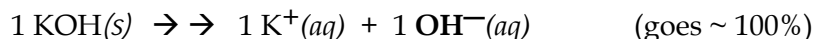
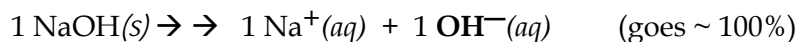
3h. $[\text{H}_3\text{O}^+] = [\text{H}^+] = 0.020 \text{ M} = 2.0 \times 10^{-2} \text{ M}$

* * * * *

Lesson 29E: Strong Base Solutions

Strong Hydroxide Bases

Strong bases, by the classical (Arrhenius) definitions, ionize essentially 100% in water to form OH^- ions. The substances most frequently used to make strong base solutions are the water soluble ionic solids sodium hydroxide (NaOH) and potassium hydroxide (KOH). In water,



There are other types of strong bases, but their reactions can be more complex. For now, we will limit our attention to the strong bases that are *alkali metal hydroxides*.

[Ions] In Strong Base Solutions

As with strong acids, if a solution of a strong base is *labeled* $[\text{NaOH}] = 0.15 \text{ M}$, this represents how the solution is *mixed*, but not the particles present in the solution. Because NaOH ionizes $\sim 100\%$ in water, what is actually present in “0.15 M NaOH” is **zero** M NaOH, 0.15 M Na^+ , and 0.15 M OH^- .

However as with strong acids, if a problem asks for the concentration of a strong base (SB) such as $[\text{NaOH}]$ or $[\text{KOH}]$, assume it is asking for not 0 M but for $[\text{SB}]_{\text{as mixed}}$.

The rules for strong bases are similar to those for strong acids.

NaOH and KOH ionize $\sim 100\%$ to form OH^- .

- For [ions], write the *REC* steps.
- If only $[\text{OH}^-]$ is needed, use the *quick rule*: $[\text{NaOH or KOH}]_{\text{mxd}} = [\text{OH}^-]_{\text{in soln.}}$
- In a *base* solution, use base rules to find $[\text{OH}^-]$, then K_W to find $[\text{H}^+]$.

We will add these to our rules that, when committed to long-term memory, simplify acid-base problem solving.

Acid-Base Fundamentals

1. See $[\text{H}^+]$ and $[\text{OH}^-]$? Write: $K_{\text{W}} = \boxed{[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}}$
2. Strong monoprotic acids ionize ~100% to form H^+ . **Alkali metal hydroxides ionize 100% to form OH^- .** For [ions], write the *REC* steps.
3. *Quick rules:* $[\text{HCl or HNO}_3]_{\text{as mixed}} = [\text{H}^+]_{\text{in soln.}}$
 $[\text{NaOH or KOH}]_{\text{as mixed}} = [\text{OH}^-]_{\text{in soln.}}$
4. See $[\text{H}_3\text{O}^+]$? Write: $\boxed{[\text{H}_3\text{O}^+] = [\text{H}^+]}$
5. In acid solutions, use acid ionization rules to find $[\text{H}^+]$, then K_{W} to find $[\text{OH}^-]$.
6. **In base solutions, use base rules to find $[\text{OH}^-]$, then K_{W} to find $[\text{H}^+]$.**

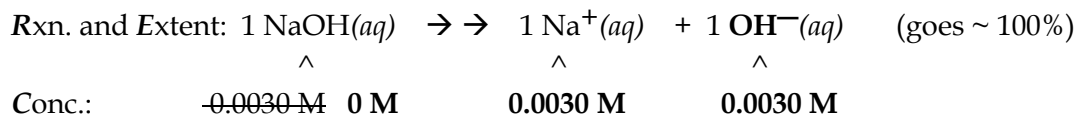
Using the new rules above, try this problem.

- Q.** In a solution labeled 0.0030 M NaOH,
- a. What three ions are present in the solution?
 - b. What is the concentration of each ion?

* * * * *

Answer

- a. NaOH ionizes to form Na^+ and OH^- ions. Water ionizes to form H^+ and OH^- .
The three ions are Na^+ , OH^- , and H^+ .
- b. For NaOH or KOH and [ions], write the *REC* steps for ~100% ionization.



which means $[\text{NaOH}]_{\text{mixed}} = \boxed{0.0030 \text{ M} = [\text{OH}^-]_{\text{in soln.}} = [\text{Na}^+]_{\text{in soln.}}}$

To find the $[\text{H}^+]$ in a base solution, first use the base rules to find $[\text{OH}^-]$ (done), then use K_{W} to find $[\text{H}^+]$.

$$K_{\text{W}} = \boxed{[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}}$$

$$[\text{H}^+] = \frac{1.0 \times 10^{-14}}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{3.0 \times 10^{-3}} = 0.33 \times 10^{-11} = \boxed{3.3 \times 10^{-12} \text{ M H}^+}$$

K_{W} check: $[\text{H}^+] \times [\text{OH}^-] = 3 \times 3.3 \approx 10$; $10^{-3} \times 10^{-12} = 10^{-15}$, combined $\approx 10 \times 10^{-15}$

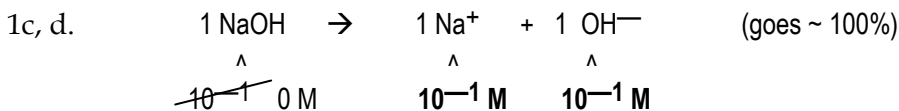
Practice: Learn the acid-base fundamental rules above, then in your notebook try these.

- In a solution labeled 10^{-1} M NaOH,
 - Write the balanced equation for the ionization of this strong base.
 - Which side is favored in this ionization: products, or reactants?
 - $[\text{OH}^-] = ?$ d. $[\text{Na}^+] = ?$ e. $[\text{H}^+] = ?$ f. $[\text{H}_3\text{O}^+] = ?$
- In a solution labeled 5.0×10^{-3} M KOH, find the
 - $[\text{OH}^-]$ b. $[\text{K}^+]$ c. $[\text{H}^+]$ d. $[\text{H}_3\text{O}^+]$
- If 5.00 millimoles NaOH are dissolved in 0.250 L soln, find
 - $[\text{OH}^-]$ b. $[\text{H}^+]$ c. $[\text{H}_3\text{O}^+]$

ANSWERS



1b. Strong bases ionize 100%, this ionization favors the **products**.



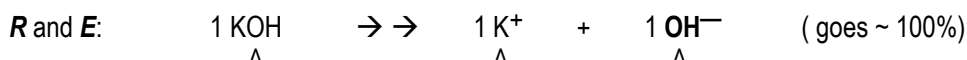
$$[\text{NaOH}]_{\text{as mixed}} = 10^{-1} \text{ M} = [\text{OH}^-]_{\text{in soln.}} = [\text{Na}^+]_{\text{in soln.}}$$

1e, f. To find $[\text{H}^+]$ in a base solution, first find $[\text{OH}^-]$ (above), then $[\text{H}^+]$ using

$$K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

$$[\text{H}^+] = \frac{1.0 \times 10^{-14}}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{10^{-1}} = 10^{-13} \text{ M} = [\text{H}^+] = [\text{H}_3\text{O}^+]$$

2a,b. For NaOH or KOH, to find [ions], write the *REC* steps for ~100% ionization.



$$[\text{KOH}]_{\text{mixed}} = \boxed{5.0 \times 10^{-3} \text{ M} = [\text{OH}^-]_{\text{in soln.}} = [\text{K}^+]_{\text{in soln.}}}$$

If we rearrange the simplified definition to solve for $[H^+]$, the result is: $[H^+] \equiv 10^{-\text{pH}}$

It may help to remember pH as the “power of H;” the number that is after the minus sign when $[H^+]$ written as a negative power of 10.

In calculations that involve pH, we will use these definitions as the

pH Prompt. If you see pH, write: $\text{pH} \equiv -\log [H^+]$ and $[H^+] \equiv 10^{-\text{pH}}$

Because our pH definitions are “shortcut” expressions of more complex equations, we will need this rule:

A pH value is not assigned a unit, but when a concentration is calculated from a pH, the unit mol/L (M) must be added to the answer.

It is also important to know that at 25 °C :

- Pure water has a pH of 7.00
- In a *pH-neutral* solution at 25°C, $[H^+] = [OH^-] = 1.0 \times 10^{-7} \text{ M}$; and $\text{pH} = 7.00$
- In *acidic* solutions, the pH is *less* than 7. In *basic* solutions, the pH is *greater* than 7.
- *Lower* pH means a *higher acidity*. Higher pH means a higher basicity.
- The further from 7 is the pH, the stronger is the acidity or basicity of the solution.
- On a base 10 logarithmic scale, changing a number by a factor of 10 changes its log by *one*.
- Since pH is a *negative log*, *increasing* $[H^+]$ by a factor of 10 *lowers* the pH by one.

Why Use the pH Scale?

The pH scale allows us to report acidity without using exponents. For non-chemists, “adjust the aquarium pH to 7.5” is easier to understand than “adjust the $[H^+]$ to 3.2×10^{-8} mol/L,” though both statements have the same meaning. For chemists, pH is a quick way to convey solution acidity.

A. Whole Number pH

In aqueous solutions, if either the $[H^+]$ or the $[OH^-]$ or the pH is known, the other two values can be calculated.

When pH values are whole numbers, pH problems do not require a calculator.

Q1. In an aqueous solution, if $[H^+] = 10^{-1} \text{ M}$, what is the pH?

★ ★ ★ ★ ★

$$[H^+] \equiv 10^{-\text{pH}} = 10^{-1} \text{ M, so pH} = 1.$$

Q2. If $[H^+] = 0.001 \text{ M}$, what is the pH? Is the solution acidic or basic?

★ ★ ★ ★ ★

$$[H^+] = 0.001 \text{ M} = 10^{-3} \text{ M} = 10^{-\text{pH}}, \text{ so pH} = 3. \text{ For pH} < 7, \text{ the solution is acidic.}$$

Q3. If the **pH = 9**, what is the **[OH⁻]**? Is the solution acidic or basic?

★ ★ ★ ★ ★

Because $[H^+] = 10^{-\text{pH}} = 10^{-9} \text{ M}$, $[OH^-] = 10^{-5} \text{ M}$ based on K_w . Note that the $[OH^-]$ is larger than the $[H^+]$, and that is the fundamental definition of a **basic** aqueous solution. The *quick* rule is: any pH greater than 7.00 is basic.

Practice A: Whole Number pH

Assume the following data apply to aqueous solutions of acids and bases. Write answers using mental or “pencil and paper” arithmetic rather than a calculator. Circle the correct answer in each *part c*.

1. Write the pH prompt from memory.

For aqueous solutions, if

- | | | | |
|---|---------------|---------------|-------------------------------------|
| 2. $[H^+] = 10^{-4} \text{ M}$, | a. $[OH^-] =$ | b. pH = | c. Is the solution acidic or basic? |
| 3. pH = 8, | a. $[H^+] =$ | b. $[OH^-] =$ | c. Is the solution acidic or basic? |
| 4. $[OH^-] = 10^{-3} \text{ M}$, | a. $[H^+] =$ | b. pH = | c. Is the solution acidic or basic? |
| 5. $[OH^-] = 1.0 \times 10^{-14} \text{ M}$, | a. $[H^+] =$ | b. pH = | c. Acidic or basic solution? |
| 6. $[H^+] = 10 \text{ M}$, | a. $[OH^-] =$ | b. pH = | c. Acidic or basic solution? |

B. From [H⁺] to Decimal pH

If the pH *or* the exponent of the hydrogen ion concentration is *not* a whole number, you can convert between $[H^+]$, $[OH^-]$, and pH using a calculator. Try this example.

Q. For an aqueous solution, if $[H^+] = 5.0 \times 10^{-4} \text{ M}$, what is the pH?

★ ★ ★ ★ ★

WANT: pH, know $[H^+]$. Write the equation that solves for pH from $[H^+]$:

$$\boxed{\text{pH} \equiv -\log [H^+]} = -\log (5.0 \times 10^{-4}) = ?$$

Solve using a calculator, round your answer to 2 decimal places, then check it below.

★ ★ ★ ★ ★

This is a *brief* review of log calculations. For detail, see Lesson 27D.

- On a standard TI-type calculator, try: 5.0 4
- An RPN calculator might use: 5.0 4
- On one type of graphing calculator: 5.0 4

Write or circle a key sequence that works.

$$\text{pH} = -\log(5.0 \times 10^{-4}) = \boxed{3.30}$$

It will simplify calculator use to take the log without the minus sign in front, then change the sign manually. For any solution with a $[\text{H}^+]$ less than 1.0 M, the pH must be *positive*. If the $[\text{H}^+]$ is greater than 1.0 M, which can occur only in relatively concentrated strong acid solutions, the pH will be a negative number.

Checking pH Calculations

Does the above answer make sense?

$$\begin{array}{ccc} \text{For } [\text{H}^+]: & 1.0 \times 10^{-4} < 5.0 \times 10^{-4}, < 10.0 \times 10^{-4} \equiv 1.0 \times 10^{-3}, \\ & \text{pH} = 4 & \text{pH} = ? & \text{pH} = 3 \end{array}$$

Based on the above, the pH should be between 4 and 3, which **3.30** is.

The estimation logic above can be used to *check* pH calculations by this rule.

pH Check: The pH rounded UP to the next whole number must equal the number after the minus sign of the exponential for the $[\text{H}^+]$ in scientific notation.

For the above problem, $\text{pH} = 3.30$ rounds *up* to **4**. The $[\text{H}^+]$ exponential term in scientific notation should therefore be 10^{-4} , and it is: 5.0×10^{-4} .

When complex operations are done on a calculator, *check* your work.

The check rule can also be used to do a quick estimate of pH from $[\text{H}^+]$.

Try **Q**. If $[\text{H}^+] = 3.0 \times 10^{-10}$ M, estimate the pH.

* * * * *

The pH must be a decimal number that rounds up to 10, which must be **9.?**

Now find a more precise answer using your calculator.

* * * * *

$$\text{pH} \equiv -\log[\text{H}^+] = -\log(3.0 \times 10^{-10}) = \mathbf{9.52} \quad \text{Check?}$$

Rounding pH Calculations

Mathematically, the statistical basis for using significant figures to convey uncertainty does not directly apply to logarithmic functions. To convey approximate uncertainty, in these lessons we will apply the following conventions to pH calculations.

- When $[\text{H}^+]$ is written as simply a power of 10, write the pH as a whole number.

Example: $[\text{H}^+] = 10^{-4}$ M, $\text{pH} = 4$;

- When $[\text{H}^+]$ is written in scientific notation, round the pH so that the number of digits in the *significand* equals the number of digits *after the decimal* in the pH.

Examples: $[\text{H}^+] = 5 \times 10^{-4}$ M, $\text{pH} = 3.3$; $[\text{H}^+] = 5.1 \times 10^{-12}$ M, $\text{pH} = 11.29$

- The K_w value on which our pH rules are based assumes a solution temperature of 25 °C, but small variations in temperature during experiments can cause significant changes in K values. Because of this and other inherent uncertainty, we will round each calculated pH to have at *most* two digits past the decimal.

Practice B: From $[H^+]$ to Decimal pH

Do the odd problems. Save the evens for your next practice session. Check answers as you go.

- Using the pH-check method above, *estimate* the pH for these questions below:
 - Problem 2: pH \approx
 - Problem 3: pH \approx
 - Problem 5: pH \approx

Use your calculator on the following.

- Find pH when $[H^+] = 2.1 \times 10^{-3}$ M.
- If $[H_3O^+] = 8.2 \times 10^{-11}$ M, what is the pH?
- If $[OH^-] = 2.0 \times 10^{-4}$ M, what is the pH?
- In a 0.040 M HCl solution, solve for the pH.
- In a 0.0030 M KOH solution, what is the pH?

C. From Decimal pH to $[H^+]$

Let's try a problem where a *non-whole-number* pH is known and the $[H^+]$ is WANTED.

Example: If the pH of a solution is 9.70, what is the $[H^+]$?

The equation that *finds* $[H^+]$ from pH is: $[H^+] \equiv 10^{-\text{pH}}$

$[H^+]$ therefore equals $10^{-9.70}$ M. That's mathematically correct, but in chemistry it is preferred to express concentrations in scientific notation, so that the exponent of 10 is a whole number. Try

- Q.** Using your calculator, convert $[H^+] = 10^{-9.70}$ M to a concentration expressed in scientific notation.

* * * * *

- On a standard TI-type calculator, try: 9.70 $\boxed{+/-}$ $\boxed{10^x}$
- A reverse Polish (RPN) calculator might use: 9.70 $\boxed{+/-}$ $\boxed{10^x}$
- On a graphing calculator, try: $\boxed{10^x}$ $\boxed{(-)}$ 9.70 $\boxed{)}$ $\boxed{\text{enter}}$ or 10 $\boxed{\wedge}$ $\boxed{(-)}$ 9.70 $\boxed{\text{enter}}$

For additional help, see Lesson 27D. Whatever method is used, write down the sequence that replicates this result:

$$[H^+] = 10^{-9.70} \text{ M} = 2.0 \times 10^{-10} \text{ M}$$

Note that whenever a concentration is calculated from a pH, the unit M must be added to the answer.

Does this answer make sense? $[H^+] = 10^{-9.70} \text{ M}$ is *close* to $[H^+] = 10^{-10} \text{ M}$. Since pH = 9.70 is a bit more acidic than pH = 10, the $[H^+]$ should be a bit higher than 10^{-10} M , and it is: $2.0 \times 10^{-10} \text{ M}$.

Q. Apply the *pH check* rule to the answer above.

* * * * *

For this problem, pH = 9.70 rounds up to 10, so the $[H^+]$ exponent in scientific notation should be 10^{-10} . The answer is 2.0×10^{-10} . Check!

Practice C: From Decimal pH to $[H^+]$

Use a calculator. Save a few problems for your next practice session.

- If pH = 5.5 a. Estimate the $[H^+]$ b. $[H^+] =$
- If pH = 8.20, a. Estimate the $[H^+]$
 - $[H^+] =$
 - $[OH^-] =$

Work Problems 3-7 in your problem notebook.

- In an HCl solution of pH = 3.60, find
 - $[H^+]$
 - $[OH^-]$
 - $[Cl^-]$
 - $[HCl]$
- If pH = 1.7, $[H^+] =$
- If pH = 7.22, $[H^+] =$
- If pH = 12.5, $[H^+] =$
- If pH = -0.50, $[H^+] =$
- Which solution in problems 3-7 is the most acidic?

D. Calculations With pOH

(If you are asked to solve problems that include pOH, complete this section.)

The pOH is the negative log of the *hydroxide* ion concentration. The pOH can be solved using the same math as pH.

In problems that include pOH, our rule will be the

pOH prompt: See pOH?

Write $pOH \equiv -\log[OH^-]$ and $[OH^-] \equiv 10^{-pOH}$ and $pH + pOH = 14.00$

The pOH is often calculated by first finding the pH and then applying the pH + pOH relationship.

The 14.00 represents the value of K_W at 25°C. Recall that “the log of a product is the sum of the logs.” If we take the log of both sides of the K_W equation:

$$\begin{aligned} \text{Since } K_W &= [\text{H}^+][\text{OH}^-] \\ \log(K_W) &= \log[\text{H}^+] + \log[\text{OH}^-] \quad \text{and} \\ -\log(K_W) &= -\log[\text{H}^+] + (-\log[\text{OH}^-]) \end{aligned}$$

And because our abbreviation for $-\log$ is p, we can write

$$\text{p}K_W = \text{pH} + \text{pOH}$$

We can use this more general relationship if we need to solve precise pH calculations at temperatures other than 25°C. Knowing the K_W value for water at a given temperature, we can write

$$\text{p}K_W = \text{pH} + \text{pOH} = -\log(K_W)$$

pH Summary: Add these to the rules for acid-base fundamentals.

7. See **pH**? Write: $\text{pH} \equiv -\log [\text{H}^+] \quad \text{and} \quad [\text{H}^+] = 10^{-\text{pH}}$

8. See **pOH**? Write: $\text{pOH} \equiv -\log[\text{OH}^-] \quad \text{and} \quad [\text{OH}^-] = 10^{-\text{pOH}}$

and $\text{pH} + \text{pOH} = 14.00 \quad (\text{at } 25^\circ\text{C}).$

Practice D: pOH Calculations

Memorize the three pOH rules above, and then solve problems 1-4 by inspection.

- In an aqueous solution, if $[\text{OH}^-] = 10^{-5} \text{ M}$,
 - $[\text{H}^+] =$
 - $\text{pH} =$
 - $\text{pOH} =$
 - Is the solution acidic or basic?
- In an aqueous solution, if $[\text{H}^+] = 10^{-4} \text{ M}$,
 - $[\text{OH}^-] =$
 - $\text{pH} =$
 - $\text{pOH} =$
 - Is the solution acidic or basic?
- In an aqueous solution, if the **pH = 12**,
 - $[\text{H}^+] =$
 - $[\text{OH}^-] =$
 - $\text{pOH} =$
 - Is the solution acidic or basic?
- In an aqueous solution, if the **pOH = 8**
 - $\text{pH} =$
 - $[\text{H}^+] =$
 - $[\text{OH}^-] =$
 - Is the solution acidic or basic?

Use a calculator for Problems 5-6. Work these in your notebook.

- 2.00 grams of NaOH is dissolved in water to make 750. mL of solution.
 - $[\text{NaOH}]_{\text{as mixed}} =$
 - $[\text{OH}^-]_{\text{in solution}} =$

4. See pH, write $\text{pH} \equiv -\log [\text{H}^+]$ and $[\text{H}^+] \equiv 10^{-\text{pH}}$

WANT pH, need $[\text{H}^+]$ to find pH, but are given $[\text{OH}^-]$.

See H^+ and OH^- ? Write: $K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$ Use K_w to find $[\text{H}^+]$:

$$[\text{H}^+] = \frac{1.0 \times 10^{-14}}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-4}} = 0.50 \times 10^{-10} = \boxed{5.0 \times 10^{-11} \text{ M H}^+}$$

(do the K_w check.)

Then $?\text{pH} \equiv -\log [\text{H}^+]$ $\text{pH} = -\log (5.0 \times 10^{-11}) = \boxed{10.30}$ (pH rounds up to 11; check.)

5. See pH, write $\text{pH} \equiv -\log [\text{H}^+]$ and $[\text{H}^+] \equiv 10^{-\text{pH}}$

WANT pH. Need $[\text{H}^+]$ first. For quick $[\text{H}^+]_{\text{in soln.}}$: $[\text{HCl}]_{\text{mixed}} = 0.040 \text{ M} = [\text{H}^+]_{\text{in soln.}}$

For pH: $\text{pH} \equiv -\log [\text{H}^+] = -\log (0.040) = -\log (4.0 \times 10^{-2}) = \boxed{1.40}$ (rounded up = 2. check.)

6. $\text{pH} \equiv -\log [\text{H}^+]$ and $[\text{H}^+] \equiv 10^{-\text{pH}}$

Need $[\text{H}^+]$ to find pH, but given is $[\text{KOH}]$. Quick: $[\text{KOH}]_{\text{mixed}} = 0.0030 \text{ M} = [\text{OH}^-]$ To find $[\text{H}^+]$,

$$[\text{H}^+] = \frac{1.0 \times 10^{-14}}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{3.0 \times 10^{-3}} = 0.33 \times 10^{-11} = \boxed{3.3 \times 10^{-12} \text{ M H}^+}$$

$\text{pH} = -\log (3.3 \times 10^{-12}) = \boxed{11.48}$ (Base solution, basic pH; 11.48 rounded up = 12. Check.)

Practice C: From Decimal pH to $[\text{H}^+]$

1. If pH = 5.5 See pH, write $\text{pH} \equiv -\log [\text{H}^+]$ and $[\text{H}^+] \equiv 10^{-\text{pH}}$

a. Estimate $[\text{H}^+]$ pH = 5.5 rounded up is 6; $[\text{H}^+]$ should be $\boxed{? \times 10^{-6} \text{ M}}$ ←compare ↑

b. The equation that finds $[\text{H}^+]$ from pH is: $[\text{H}^+] = 10^{-\text{pH}}$; $[\text{H}^+] = 10^{-5.5} = \boxed{3 \times 10^{-6} \text{ M}}$

2. See pH? Write: $\text{pH} \equiv -\log [\text{H}^+]$ and $[\text{H}^+] \equiv 10^{-\text{pH}}$

2a. Estimate: pH = 8.20 rounds up to 9, $[\text{H}^+]$ should = $?.? \times 10^{-9} \text{ M}$

WANT $[\text{H}^+]$ from pH, use $[\text{H}^+] = 10^{-\text{pH}}$; $[\text{H}^+] = 10^{-8.20} = \boxed{6.3 \times 10^{-9} \text{ M H}^+}$ Checks vs. pH.

WANT $[\text{OH}^-]$. Know $[\text{H}^+]$ and pH. K_w relates $[\text{H}^+]$ and $[\text{OH}^-]$: $[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$;

$$? = [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{[\text{H}^+]} = \frac{1.0 \times 10^{-14}}{6.3 \times 10^{-9}} = 0.16 \times 10^{-5} = \boxed{1.6 \times 10^{-6} \text{ M OH}^-}$$

3. See pH? Write $\text{pH} \equiv -\log [\text{H}^+]$ and $[\text{H}^+] \equiv 10^{-\text{pH}}$

a. $[\text{H}^+] = ?$ $[\text{H}^+] \equiv 10^{-\text{pH}} = 10^{-3.60} = 2.5 \times 10^{-4} \text{ M}$

b. $[\text{OH}^-] = ?$ Knowing $[\text{H}^+]$, to find $[\text{OH}^-]$, use K_w .

$$? = [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{[\text{H}^+]} = \frac{1.0 \times 10^{-14}}{2.5 \times 10^{-4}} = 0.40 \times 10^{-10} = 4.0 \times 10^{-11} \text{ M OH}^-$$

c, d. $[\text{HCl}]_{\text{mixed}} = [\text{H}^+]_{\text{soln}} = [\text{Cl}^-]_{\text{soln}}$ From Part a, $[\text{H}^+] = 2.5 \times 10^{-4} \text{ M} = [\text{Cl}^-] = [\text{HCl}]$

4. If $\text{pH} = 1.7$, $[\text{H}^+] \equiv 10^{-\text{pH}} = 10^{-1.7} = 2 \times 10^{-2} \text{ M}$

5. If $\text{pH} = 7.22$, $[\text{H}^+] \equiv 10^{-\text{pH}} = 10^{-7.22} = 6.0 \times 10^{-8} \text{ M}$

6. $\text{pH} = 12.5$, $[\text{H}^+] \equiv 10^{-\text{pH}} = 10^{-12.5} = 3 \times 10^{-13} \text{ M}$

7. $\text{pH} = -0.50$, $[\text{H}^+] \equiv 10^{-\text{pH}} = 10^{-(-0.50)} = 10^{0.50} = 3.2 \text{ M}$

Solutions with a $[\text{H}^+]$ higher than 1.0 M have a negative pH.

8. The Problem 7 solution has the both the lowest pH and the highest $[\text{H}^+]$. By either measure it is the most acidic.

Practice D: pOH Calculations

1. If $[\text{OH}^-] = 10^{-5}$, a. $[\text{H}^+] = 10^{-9} \text{ M}$ b. $\text{pH} = 9$ c. $\text{pOH} = 5$ d. **Basic**

2. If $[\text{H}^+] = 10^{-4}$, a. $[\text{OH}^-] = 10^{-10} \text{ M}$ b. $\text{pH} = 4$ c. $\text{pOH} = 10$ d. **Acidic**

3. If $\text{pH} = 12$, a. $[\text{H}^+] = 10^{-12} \text{ M}$ b. $[\text{OH}^-] = 10^{-2} \text{ M}$ c. $\text{pOH} = 2$ d. **Basic**

4. If $\text{pOH} = 8$ a. $\text{pH} = 6$ b. $[\text{H}^+] = 10^{-6} \text{ M}$ c. $[\text{OH}^-] = 10^{-8} \text{ M}$ d. **Acidic**

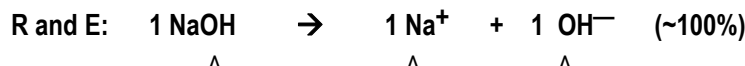
5. a. WANT: ? mol NaOH = $[\text{NaOH}]_{\text{mixed}}$
L soln

DATA: 2.00 grams NaOH = 750. mL soln (equivalent: two measures of same solution.)

40.0 grams NaOH = 1 mole NaOH (grams prompt)

SOLVE: ? moles NaOH = $\frac{2.00 \text{ g NaOH}}{750 \text{ mL soln}} \bullet \frac{1 \text{ mol NaOH}}{40.0 \text{ g NaOH}} \bullet \frac{1 \text{ mL}}{10^{-3} \text{ L}} = 0.0667 \text{ M NaOH}$

b. See NaOH and [ions]? Write the REC steps: Reaction, Extent, Concentrations.



$[\text{NaOH}]_{\text{mixed}} = 0.0667 \text{ M} = [\text{OH}^-]_{\text{in soln.}} = [\text{Na}^+]_{\text{in soln.}}$

b. $[\text{OH}^-] = ?$ $[\text{NaOH}]_{\text{mxd.}} = 0.0667 \text{ M (part a)} = [\text{OH}^-]_{\text{in soln}} = [\text{Na}^+]_{\text{in soln}} = \text{Part c answer}$

d. $[H^+]$ WANT $[H^+]$, know $[OH^-]$.

$$[H^+] = \frac{1.0 \times 10^{-14}}{[OH^-]} = \frac{1.0 \times 10^{-14}}{6.67 \times 10^{-2}} = 0.15 \times 10^{-12} = 1.5 \times 10^{-13} \text{ M } H^+$$

e. $pOH = ?$

See pOH ? Write: $pOH \equiv -\log[OH^-]$ and $[OH^-] \equiv 10^{-pOH}$ and $pH + pOH = 14.00$

$$pOH \equiv -\log[OH^-] = -\log(0.0667) = -\log(6.67 \times 10^{-2}) = 1.18 = pOH$$

f. pH ? Since $pH + pOH = 14.00$,

$$pH = 14.00 - pOH = 14.00 - 1.18 = 12.82 \quad (\text{compare } pH \text{ to } [H^+] \text{ above.})$$

6. The pOH in a KOH solution is 1.5. What is the $[KOH]$?

b. See KOH and [particles]? Write the *REC* steps. See pOH ? Use the pOH prompt.



$$[OH^-] \equiv 10^{-pOH} = 10^{-1.5} \text{ M} = 3 \times 10^{-2} \text{ M} = [OH^-] = X \text{ M} = [KOH]_{\text{mixed}}$$

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Summary -- Acid-Base Fundamentals

If you have not already done so, you may want to design flashcards that will assist in adding these rules to your memory.

Acid-Base Fundamentals

- See $[H^+]$ and $[OH^-]$? Write: $K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$
- In water,
 - strong monoprotic acids ionize ~100% and form H^+ .
 - Alkali metal hydroxides ionize 100% and form OH^- .
 - For [ions], write the *REC* steps or use the quick rules.
- Quick rules: $[HCl \text{ or } HNO_3]_{\text{as mixed}} = [H^+]_{\text{in soln.}}$
 $[NaOH \text{ or } KOH]_{\text{mixed}} = [OH^-]_{\text{in soln.}}$
- See $[H_3O^+]$? Write: $[H_3O^+] = [H^+]$
- In acid solutions, use acid ionization rules to find $[H^+]$, then K_w to find $[OH^-]$.
- In *base* solutions, use base rules to find $[OH^-]$, then K_w to find $[H^+]$.
- See pH? Write: $pH \equiv -\log [H^+]$ and $[H^+] = 10^{-pH}$
- See pOH? Write: $pOH \equiv -\log [OH^-]$ and $[OH^-] = 10^{-pOH}$
 and $pH + pOH = 14.00$ (at 25°C)
- pH values are not given units. When a concentration is calculated based on pH, the unit mol/L (M) must be assigned to the answer.
- Using the pH scale, for aqueous solutions at 25°C:
 - Pure water has a pH of 7.00
 - In a *neutral* solution, $[H^+] = [OH^-] = 1.0 \times 10^{-7} \text{ M}$; and $pH = 7.00$
 - In *acidic* solutions, pH is *less* than 7. In *basic* solutions, pH is *greater* than 7.
 - A lower pH means a higher acidity. A higher pH means lower acidity and a higher basicity.
 - The further from 7 is pH, the stronger is the acidity or basicity of the solution.
- The pH check:* The pH rounded UP to the next whole number must equal the number after the minus sign of the exponential for the $[H^+]$ in scientific notation.

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