# CHAPTER 16: ACIDS AND BASES

## Enhanced Introductory College Chemistry

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## Chapter Contents

Please visit the web version of Enhanced Introductory College Chemistry (https://ecampusontario.pressbooks.pub/enhancedchemistry/) to access the complete book, interactive activities and ancillary resources.

## In this chapter, you will learn about

- General characteristics and properties of acids and bases.

- Two models of acids and bases and the relationship between conjugate acid-base pairs.
- The autoionization of water.
- pH, pOH and the pH scale.
- How to calculate pH, pOH, and the acid and base concentration of various solutions.
- Neutralization reactions and how to do calculations involving strong acids and strong bases.
- Acid-Base titrations.
- The general characteristics of buffered solutions.

To better support your learning, you should be familiar with the following concepts before starting this chapter:

- How to use your scientific calculator including the following functions (buttons): log, $10^x$, scientific notation.
- Review double displacement reactions and net ionic equations.
- Understand concepts relating to ionization/dissociation of salts in solution.
- Polarity characteristics of water and its molecular equation.
- Performing molarity calculations of solutions to determine a solution's concentration.

**Figure 16a** Sinkhole Shown in Nature: Cenote Angelita is located in the state of Quintana Roo, in the Yucatan Peninsula of Mexico. Sinkholes in nature, such as this one, are the result of reactions between acidic groundwaters and basic rock formations, like limestone. (credit: work by Offthebeatonpath, CC BY-SA 3.0)

In our bodies, in our homes, and in our industrial society, acids and bases play key roles. Proteins, enzymes, blood, genetic material, and other components of living matter contain both acids and bases. We seem to like the sour taste of acids; we add them to soft drinks, salad dressings, and spices. Many foods, including citrus fruits and some vegetables, contain acids. Cleaners in our homes contain acids or bases. Acids and bases not only play important roles in the health/medical industry but also in the chemical industry. Huge quantities of sulfuric acid ($H_2SO_4$ ), ammonia ($NH_3$), urea ($CH_4N_2O$), and phosphoric acid ($H_3PO_4$) are released into the atmosphere, globally, every year. In Canada, the production of sulfuric acid is used by industry in the production of phosphate fertilizers, bleaching agents in pulp and paper manufacturing, and waste water processing in sewage treatment plants (Environment and Climate Change Canada, 2022, Sources and uses of sulfuric acid, para. 1). According to National Pollutant Release Inventory Overview by Environment and Climate Change Canada (2022),

> “the primary sources of sulfuric acid emissions are manufacturing, coal-fired power plants, petroleum and coal product refining and non-conventional oil extraction (including oil sands). These emissions are mostly to the air, with total air releases of 3485 tonnes in 2020. Electric power facilities, the manufacturing industry and pulp and paper plants discharged a combined 71 tonnes to water in 2020. A combined 16 tonnes of sulfuric acid was released to land from the mining and manufacturing industries. When burned, the sulfur content in fossil fuels is mostly converted into sulfur dioxide ($SO_2$), which can further oxidize into sulfur trioxide ($SO_3$) and then react with water to form sulfuric acid” (Releases of sulfuric acid section, para. 1).

For more information about sulfuric acid release in Canada or other pollution and waste management inventories visit the National Pollutant Release Inventory: tools and resources (https://www.canada.ca/en/environment-climate-change/services/national-pollutant-release-inventory/tools-resources-data.html)

Another factor that plays a crucial role in influencing chemical changes in our environment is acid rain. When precipitation from rain, snow, sleet, or hail falls to the earth, it collects acidic particles and gases dissolved in the atmosphere and turns the precipitation more acidic. Not only does acid deposition damage various ecosystems and physical structures, but also has negative health effects on its inhabitants, such as fish, wildlife, and human populations, (Environment And Climate Change Canada, 2018). To reference more about acid rain and its effects in Canada link visit Acid rain: causes and effects (https://www.canada.ca/en/environment-climate-change/services/air-pollution/issues/acid-rain-causes-effects.html).

The history of the acid rain issues and its effects on the environment is a very intriguing phenomenon. **Watch Whatever happened to acid rain (5:39). (https://www.youtube.com/watch?v=WIdWjqZsGgg)**

This chapter will illustrate the chemistry of acid-base reactions and equilibria, and provide you with tools for quantifying the concentrations of acids and bases in solutions.

## Attribution & References

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References:

- Environment And Climate Change Canada. (2022, August 17). *National Pollutant Release Inventory overview: Sulphuric acid.* Canada.ca. Retrieved December 6, 2022.
- Environment And Climate Change Canada. (2018, June 27). *Acid rain: Causes and effects.* Canada.ca. Retrieved December 6, 2022.

# 16.1 ACIDS AND BASES

### Learning Objectives

By the end of this section, you will be able to:

- Examine properties of acids and bases and provide examples of both
- Define a strong and a weak acid and base. Using a table for reference, recognize an acid or a base as strong or weak.
- Define an Arrhenius acid and Arrhenius base
- Identify acids, bases, and conjugate acid-base pairs according to the Brønsted-Lowry definition
- Explain the Lewis model of acid-base chemistry

Many people enjoy drinking coffee. A cup first thing in the morning helps start the day. But keeping the coffee maker clean can be a problem. Lime deposits build up after a while and slow down the brewing process. The best cure for this is to put vinegar (dilute acetic acid) in the pot and run it through the brewing cycle. The vinegar dissolves the deposits and cleans the maker, which will speed up the brewing process back to its original rate. Just be sure to run water through the brewing process after the vinegar, or you will get some really horrible tasting coffee. Acids and bases are common substances found in many every day items, from fruit juices and soft drinks to soap. In this section, we'll explore the basic properties of acids and bases, and learn about the chemical nature of these important compounds.

**Watch The Strengths and Weaknesses of Acids and Bases (3min 47s). (https://youtu.be/DupXDD87oHc)**

## Acids: Properties and Examples

Acids are very common in some of the foods that we eat. Citrus fruits such as oranges and lemons contain citric acid and ascorbic acid, which is better known as vitamin C. Carbonated sodas contain phosphoric acid. Vinegar contains acetic acid. Your own stomach utilizes hydrochloric acid to digest food. Acids are a distinct class of compounds because of the properties of their aqueous solutions as outlined below:

1. Aqueous solutions of acids are electrolytes, meaning that they conduct electrical current. Some acids are strong electrolytes because they ionize completely in water, yielding a great many ions. Other acids are weak electrolytes that exist primarily in a non-ionized form when dissolved in water.
2. Acids have a sour taste. Lemons, vinegar, and sour candies all contain acids.
3. Acids change the colour of certain acid-base indicates. Two common indicators are litmus and phenolphthalein. Blue litmus turns red in the presence of an acid, while phenolphthalein turns colourless.
4. Acids react with bases to produce a salt compound and water. When equal moles of an acid and a base are combined, the acid is neutralized by the base. The products of this reaction are an ionic compound, which is labeled as a salt, and water.
5. Acids react with active metals to yield hydrogen gas. Recall that an activity series is a list of metals in descending order of reactivity. Metals that are above hydrogen in the activity series will replace the hydrogen from an acid in a single-replacement reaction (Chapter 8.3), as shown below:

$$\mathrm{Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)}$$

Table 16.1a provides a partial list of some common acids and their uses.

**Table 16.1a Common Acids and Their Uses**

| **Compound Name and Chemical Formula** | **Common Name (if applicable)** | **Uses** |
|---|---|---|
| hydrochloric acid, $HCl$ | muriatic acid (used in pools) and stomach acid is HCl | Used in cleaning (refining) metals, in maintenance of swimming pools, and for household cleaning. |
| sulfuric acid, $H_2SO_4$ | none | Used in car batteries, and in the manufacture of fertilizers. |
| nitric acid, $HNO_3$ | none | Used in the manufacture of fertilizers, explosives and in extraction of gold. |
| acetic acid, $HC_2H_3O_2$ | vinegar | Main ingredient in vinegar. |
| carbonic acid, $H_2CO_3$ | responsible for the "fizz" in carbonated drinks | As an ingredient in carbonated drinks. |
| citric acid, $C_6H_8O_7$ | none | Used in food and dietary supplements. Also added as an acidulant in creams, gels, liquids, and lotions. |
| acetylsalicylic acid, $C_6H_4(OCOCH_3)CO_2H$ | aspirin | The active ingredient in aspirin. |

Later in this chapter, we will examine the chemistry behind what exactly makes an acid behave as an acid and what makes a base behave as a base. For now, take a look at the formulas given in the table 16.1a and take a guess as to what common feature characterizes an acid.

**Figure 16.1a** An Image of a solution containing Phenolphthalein Indicator: Phenolphthalein is often used as an indicator in acid–base titrations. It is colourless in acidic solutions and typically pink in basic solutions between a pH of 8.5-10. (credit: work by Ben Mills, PD)

## Bases: Properties and Examples

Perhaps you have eaten too much pizza and felt very uncomfortable hours later. This feeling is due to excess stomach acid being produced. The discomfort can be dealt with by taking an antacid. The base in the antacid will react with the HCl in the stomach and neutralize it, taking care of that unpleasant feeling. Bases have properties that mostly contrast with those of acids.

1. Aqueous solutions of bases are also electrolytes. Bases can be either strong or weak, just as acids can.
2. Bases often have a bitter taste and are found in foods less frequently than acids. Many bases, like soaps, are slippery to the touch.
3. Bases also change the colour of indicators. Litmus turns blue in the presence of a base, while phenolphthalein turns pink as shown in Figure 16.1a
4. Bases do not react with metals in the way that acids do but tend to react with fats and oils
5. Bases react with acids to produce a salt and water.

## Warning!

Tasting chemicals and touching them are NOT good lab practices and should be avoided—in other words—don't do this at home.

Bases are less common as foods, but they are nonetheless present in many household products. Many cleaners contain ammonia, a base. Sodium hydroxide is found in drain cleaner. Antacids, which combat excess stomach acid, are comprised of bases such as magnesium hydroxide or sodium hydrogen carbonate.

Various common bases and corresponding uses are given in Table 16.1b.

**Table 16.1b Common Bases and Corresponding Uses**

| **Compound Name and Chemical Formula** | **Common Name (if applicable)** | **Uses** |
|---|---|---|
| sodium hydroxide, NaOH | (lye or caustic soda) | Used in the manufacture of soaps and detergents, and as the main ingredient in oven and drain cleaners. |
| potassium hydroxide, KOH | (lye or caustic potash) | Used in the production of liquid soaps and soft soaps. Used in alkaline batteries. |
| magnesium hydroxide, $Mg(OH)_2$ | (milk of magnesia) | Used as an ingredient in laxatives, antacids, and deodorants. Also used in the neutralization of acidic wastewater. |
| calcium hydroxide, $Ca(OH)_2$ | (slaked lime) | Used in the manufacture of cement and lime water. Also, added to neutralize acidic soil. |
| aluminum hydroxide, $Al(OH)_3$ | | Used in water purification and as an ingredient in antacids. |
| ammonia, $NH_3$ | | Used as a building block for the synthesis of many pharmaceutical products and in many commercial cleaning products. Used in the manufacture of fertilizers. |

To learn more about commercially available acids and bases, see Appendix G.

## Exercise 16.1a

Check Your Learning Exercise (Text Version)
Question 1 of Interactive:

For each of the following descriptions or properties, determine whether it is describing an acid or base:

1. Typically have a bitter taste
2. Solutions feel slippery
3. Turns litmus paper red
4. Turns litmus paper blue
5. Typically react with metals, producing hydrogen bubbles (gas)
6. Type of substance that aids in protein digestion in the stomach
7. The active ingredient of aspirin is this type of substance
8. Typically have a sour taste
9. Baking soda is classified as this
10. Citrus fruits are classified as this
11. Antacids are classified as this
12. Vinegar is classified as this

Question 2 of Interactive:

All of the follow are TRUE statements about bases EXCEPT

1. often an ingredient in cleaning solutions and feel slippery
2. when dissolved in water, have a bitter taste
3. Are electrolytes – conduct electricity
4. when metals are placed in a basic solution, they react, producing hydrogen bubbles

Question 3 of Interactive:

All of the follow are TRUE statements about ACIDS EXCEPT

1. acids are found in citrus fruits and provide the bitter taste
2. acids are found in citrus fruits and provide the sour taste
3. turns litmus paper red
4. Are electrolytes – conduct electricity

Question 4 of Interactive:

True or False? Both strong acids and bases can ionize in solution, which makes them electrolytes and can conduct an electrical current. However, only bases can dissolve metals, producing hydrogen gas.

### Check Your Answer[1]

## Strength of Acids and Bases

An acid or base's strength refers to its degree of ionization. We will explore the basics of acid/base strength below.

### Strong and Weak Acids

Except for their names and formulas, so far we have treated all acids as equals, especially in a chemical reaction. However, acids can be very different in a very important way. Consider HCl(*aq*). When HCl is dissolved in water ($H_2O$), it completely dissociates (100%) into $H^+(aq)$ and $Cl^-(aq)$ ions; all the HCl molecules become ions:

$$HCl \rightarrow H^+(aq) + Cl^-(aq)$$

Any acid that dissociates 100% into ions is called a **strong acid**. If it does not dissociate 100%, it is a **weak acid**. Acetic Acid, $HC_2H_3O_2$, is an example of a weak acid (it does not dissociate completely in aqueous solution, only about 5%):

$$HC_2H_3O_2 \rightarrow H^+(aq) + C_2H_3O^{2-}(aq)$$

Because this reaction does not go 100% to completion, it is more appropriate to write it as a reversible reaction:

$$HC_2H_3O_2 \rightleftharpoons H^+(aq) + C_2H_3O^{2-}(aq)$$

As it turns out, there are very few strong acids, which are given in Table 16.1c. If an acid is not listed here, it is a weak acid. It may be 1% ionized or 99% ionized, but it is still classified as a weak acid.

Any acid that dissociates 100% into ions is called a strong acid. If it does not dissociate 100%, it is a weak acid.

**Table 16.1c Strong Acids and Bases**

| Acids | Bases |
|---|---|
| HCl | LiOH |
| HBr | NaOH |
| HI | KOH |
| $HNO_3$ | RbOH |
| $H_2SO_4$ | CsOH |
| $HClO_3$ | $Mg(OH)_2$ |
| $HClO_4$ | $Ca(OH)_2$ |
| [Blank] | $Sr(OH)_2$ |
| [Blank] | $Ba(OH)_2$ |

## Strong and Weak Bases

The issue is similar with bases: a **strong base** is a base that is 100% ionized in solution. If it is less than 100% ionized in solution, it is a **weak base**. There are very few strong bases (Table 16.1c); any base not listed is a weak base. All strong bases are compounds containing hydroxide, $OH^-$. Any other molecules classified as a base (using other mechanisms on defining bases), such as $NH_3$ (which does not contain $OH^-$ ions as part of its formula), will be a weak base.

# Molecular Definitions of Acids and Bases

Acids and bases have been known for a long time. When Robert Boyle characterized them in 1680, he noted that acids dissolve many substances, change the colour of certain natural dyes (for example, they change litmus from blue to red), and lose these characteristic properties after coming into contact with alkalis (bases). In the eighteenth century, it was recognized that acids have a sour taste, react with limestone to liberate a gaseous substance (now known to be $CO_2$), and interact with alkalis to form neutral substances. In 1815, Humphry Davy contributed greatly to the development of the modern acid-base concept by demonstrating that hydrogen is the essential constituent of acids. Around that same time, Joseph Louis Gay-Lussac concluded that acids are substances that can neutralize bases and that these two classes of substances can be defined only in terms of each other. Three main classifications of acids and bases are predominantly used to explain acid base reactions:

1. Arrhenius acids and bases: Acids yield protons when dissolved in solution, while Arrhenius bases yields hydroxide ions.

2. Brønsted-Lowry acids and bases. Acids are proton donors, while Brønsted-Lowry bases are proton acceptors.
3. Lewis acids and bases: Acids are electron acceptors, while Lewis bases are electron donors.

## Arrhenius Acids/Bases

The significance of hydrogen was reemphasized in 1884 when Carl Axel Arrhenius defined an acid as a compound that dissolves in water to yield hydrogen cations (now recognized to be hydronium ions, $H_3O^+$) and a base as a compound that dissolves in water to yield hydroxide anions ($OH^-$). In an earlier chapter on chemical reactions, we defined acids and bases as Arrhenius did:

- An acid (**Arrhenius acid**) as a compound that dissolves in water to produce $H^+$ ions (in other words, yields hydronium ions, $H_3O^+$).
- A base (**Arrhenius base**) as a compound that dissolves in water to to yield hydroxide ions ($OH^-$).

This definition is not wrong; it is simply limited. The theory does not explain the weak base ammonia ($NH_3$), which in the presence of water, releases hydroxide ions into solution, but does not contain $OH^-$ itself. The Arrhenius definition of acid and base is also limited to aqueous solutions.

## Brønsted-Lowry Acids/Bases

Later in 1923 the Danish chemist Johannes Brønsted and the English chemist Thomas Lowry proposed a more general definition of an acid or a base. Their definition centres on the proton, $H^+$. A proton is what remains when a normal hydrogen atom, ${}_1^1\text{H}$, loses an electron. A compound that donates a proton to another compound is called a **Brønsted-Lowry acid**, and a compound that accepts a proton is called a **Brønsted-Lowry base**. An acid-base reaction is the transfer of a proton from a proton donor (acid) to a proton acceptor (base).

Acids may be compounds such as HCl or $H_2SO_4$, organic acids like acetic acid ($CH_3COOH$) or ascorbic acid (vitamin C), or $H_2O$. Anions (such as $HSO_4^-$, $H_2PO_4^-$ $HS^-$, and $HCO_3^-$ and cations (such as $H_3O^+$, $NH_4^+$, and $[Al(H_2O)_5OH]^{2+}$) may also act as acids. Bases fall into the same three categories. Bases may be neutral molecules (such as $H_2O$, $NH_3$, and $CH_3NH_2$), anions (such as $OH^-$, $HS^-$, $HCO_3^-$, $CO_3^{2-}$, $F^-$, and $PO_4^{3-}$), or cations (such as $[Al(H_2O)_5OH]^{2+}$). The most familiar bases are ionic compounds such as NaOH and $Ca(OH)_2$, which contain the hydroxide ion, $OH^-$. The hydroxide ion in these compounds accepts a proton from acids to form water:

$$H^+ + OH^- \longrightarrow H_2O$$

We call the product that remains after an acid donates a proton the **conjugate base** of the acid. This species is a base because it can accept a proton (to re-form the acid):

$$\begin{aligned}
\text{acid} &\leftrightharpoons \text{proton} + \text{conjugate base} \\
\text{HF} &\leftrightharpoons \text{H}^+ + \text{F}^- \\
\text{H}_2\text{SO}_4 &\leftrightharpoons \text{H}^+ + \text{HSO}_4{}^- \\
\text{H}_2\text{O} &\leftrightharpoons \text{H}^+ + \text{OH}^- \\
\text{HSO}_4{}^- &\leftrightharpoons \text{H}^+ + \text{SO}_4{}^{2-} \\
\text{NH}_4{}^+ &\leftrightharpoons \text{H}^+ + \text{NH}_3
\end{aligned}$$

## Example 16.1a

Write the conjugate base for each of the following acids: (Hint! Remember an acid will donate a proton)

a. $HClO_4$
b. $H_3PO_4$
c. $CH_3NH_3^+$

### Solution

a. $HClO_4$ is the acid; $ClO_4^-$ is its conjugate base
b. $H_3PO_4$ is the acid; $H_2PO_4^-$ is its conjugate base
c. $CH_3NH_3^+$ is the acid; $CH_3NH_2$ is its conjugate base

**Source:** "Example 16.1a" by Jackie MacDonald is licensed under CC BY-NC 4.0.

## Exercise 16.1b

Write the conjugate base for each of the following acids:

a. $NH_2^-$

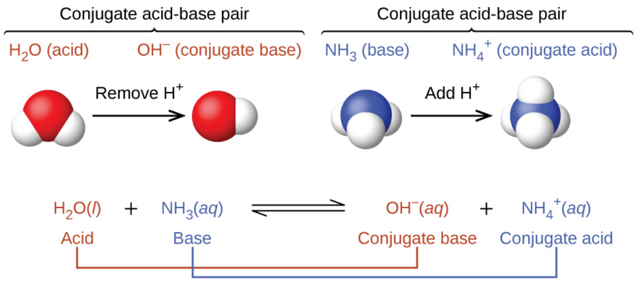
b. HBr
c. $HSO_4^-$

**Check Your Answer**[2] **Source: “Exercise 16.1b” by Jackie MacDonald is licensed under CC BY-NC 4.0.**

We call the product that results when a base accepts a proton the base’s **conjugate acid**. This species is an acid because it can give up a proton (and thus re-form the base):

$$\begin{aligned}
\text{base} + \ \text{proton} &\leftrightharpoons \text{conjugate acid} \\
OH^- \ + \ H^+ &\leftrightharpoons H_2O \\
H_2O \ + \ H^+ &\leftrightharpoons H_3O^+ \\
NH_3 \ + \ H^+ &\leftrightharpoons {NH_4}^+ \\
S^{2-} \ + \ H^+ &\leftrightharpoons HS^- \\
{CO_3}^{2-} \ + \ H^+ &\leftrightharpoons {HCO_3}^- \\
F^- \ + \ H^+ &\leftrightharpoons HF
\end{aligned}$$

In these two sets of equations, the behaviours of acids as proton donors and bases as proton acceptors are represented in isolation. In reality, all acid-base reactions involve the *transfer* of protons between acids and bases. Interestingly, water can act as an acid or base depending on what gets dissolved in it. For example, consider the acid-base reaction that takes place when ammonia is dissolved in water (Figure 16.1b). A water molecule (functioning as an acid) transfers a proton to an ammonia molecule (functioning as a base), yielding the conjugate base of water, $OH^-$, and the conjugate acid of ammonia, $NH_4^+$.



**16.1b** Acid-Base Reaction between Water and Ammonia showing Conjugate Acid-Base Pairs: A Brønsted-Lowry acid reacts with a base to form the conjugate base and acid, as shown with $H_2O$ (acid) reacting with $NH_3$ (base) to form their respective conjugates. (credit: *Chemistry 2e (OpenStax)*, CC-BY-4.0).

The reaction between a Brønsted-Lowry acid and water is called acid ionization. For example, when hydrogen fluoride dissolves in water and ionizes, protons are transferred from hydrogen fluoride molecules to water molecules, yielding hydronium ions and fluoride ions:

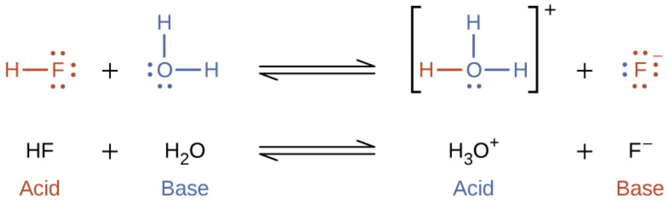


**Figure 16.1c** The Acid Ionization Reaction of an Acid (HF) with Water: The acid ionization reaction of an acid (HF) with water, resulting in formation of hydronium ions and the conjugate base of the acid. In this reaction, the water molecule functions as an base since it accepts a proton from the hydrofluoric acid (HF) molecule (functioning as a Brønsted-Lowry acid), yielding the conjugate acid of water, a hydronium ion ($H_3O^+$) and the conjugate base of hydrofluoric acid, $F^-$. (credit: *Chemistry 2e (OpenStax)*, CC-BY-4.0).

When we add a base to water, a base ionization reaction occurs in which protons are transferred from water molecules to base molecules. For example, adding pyridine to water yields hydroxide ions and pyridinium ions:

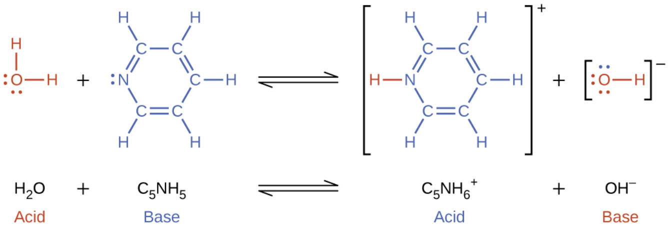


**Figure 16.1d** Pyridine Reacts with Water to Form the Pyridinium Ion and Hydronium Ion: In this reaction, the water molecule functions as an acid since it donates a proton to pyridine (functioning as a Brønsted-Lowry base), yielding the conjugate base of water, hydroxide ions ($OH^-$) and the conjugate acid of pyridine, pyridinium ions ($C_5NH_6^+$). (credit: *Chemistry 2e (OpenStax)*, CC-BY-4.0).

Notice that both these ionization reactions in Figure 16.1c and 16.1d are represented as equilibrium processes. The relative extent to which these acid and base ionization reactions proceed is an important topic that will be covered in the equilibrium chapter.

**Watch Conjugate Acids & Bases | Acids, Bases & Alkali's | Chemistry | FuseSchool (3min 45sec) (https://youtu.be/fUhPLe0NAvA).**

## Example 16.1b

Identify and label the Brønsted-Lowry acid, its conjugate base, the Brønsted-Lowry base, and its conjugate acid in each of the following equations:

a. $NO^{2-} + H_2O \rightarrow HNO_2 + OH^-$
b. $HBr + H_2O \rightarrow H_3O^+ + Br^-$

### Solution

a. $H_2O$ is the Brønsted-Lowry acid and $OH^-$ is its conjugate base; $NO^{2-}$ is the Brønsted-Lowry base and $HNO_2$ is its conjugate acid.
b. HBr is the Brønsted-Lowry acid and $Br^-$ is its conjugate base; $H_2O$ is the Brønsted-Lowry base and $H_3O^+$ is its conjugate acid.

## Exercise 16.1c

Identify and label the Brønsted-Lowry acid, its conjugate base, the Brønsted-Lowry base, and its conjugate acid in each of the following equations:

a. $HS^- + H_2O \rightarrow H_2S + OH^-$
b. $H_2PO_4^- + OH^- \rightarrow HPO_4^{2-} + H_2O$
c. $H_2PO_4^- + HCl \rightarrow H_3PO_4 + Cl^-$
d. $[Fe(H_2O)_5(OH)]^{2+} + [Al(H_2O)_6]^{3+} \rightarrow [Fe(H_2O)_6]^{3+} + [Al(H_2O)_5(OH)]^{2+}$

**Check Your Answer**[3]

## Exercise 16.1d

Check Your Learning Exercise (Text Version)

1. Correctly match each term with its correct definition:
   TERMS: (i) Brønsted-Lowry acid; (ii) Brønsted-Lowry base
   DEFINITIONS: (1) A compound that donates a proton to another compound; (2) a compound that accepts a proton
2. What is the conjugate acid in the following equation? $PO_4^{3-} + HNO_3 \rightarrow NO_3^- + HPO_4^{2-}$
3. What is the conjugate base in the following equation? $HCO_3^- + HCl \rightarrow H_2CO_3 + Cl^-$
4. Which of the following is the correct Brønsted-Lowry acid, conjugate base pair for the following reaction: $CH_3OH + H^- \rightarrow CH_3O^- + H_2$

    1. Brønsted-Lowry acid $CH_3O^-$, conjugate base $CH_3OH$
    2. Brønsted-Lowry acid $CH_3OH$, conjugate base $CH_3O^-$
    3. Brønsted-Lowry acid $H^-$ , conjugate base $H_2$
    4. Brønsted-Lowry acid $CH_3OH$, conjugate base $H^-$

### Check Your Answer[4]

**Source:** "Exercise 16.1d" by Jackie MacDonald, licensed under CC BY-NC 4.0.

## Lewis Acids and Bases

In 1923, G. N. Lewis proposed a generalized definition of acid-base behaviour in which acids and bases are identified by their ability to accept or to donate a pair of electrons and form a coordinate covalent bond.

A **coordinate covalent bond** (or dative bond) occurs when one of the atoms in the bond provides both bonding electrons. For example, a coordinate covalent bond occurs when a water molecule combines with a hydrogen ion to form a hydronium ion. A coordinate covalent bond also results when an ammonia molecule combines with a hydrogen ion to form an ammonium ion. Both of these equations are shown in Figure 16.1e.

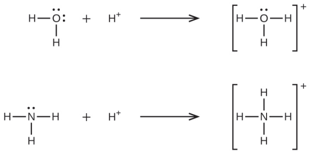


**Figure 16.1e** Illustrating Coordinate Covalent Bonds: Two chemical equations are shown to illustrate Lewis acids and base reactions: the formation of a hydronium ion and an ammonium ion. (credit: *Chemistry (OpenStax)*, CC BY 4.0)

A **Lewis acid** is any species (molecule or ion) that can accept a pair of electrons, and a **Lewis base** is any species (molecule or ion) that can donate a pair of electrons.

A Lewis acid-base reaction occurs when a base donates a pair of electrons to an acid. A **Lewis acid-base adduct**, a compound that contains a coordinate covalent bond between the Lewis acid and the Lewis base, is formed. The following equations illustrate the general application of the Lewis concept.

The boron atom in boron trifluoride, $BF_3$, has only six electrons in its valence shell. Being short of the preferred octet, $BF_3$ is a very good Lewis acid and reacts with many Lewis bases; a fluoride ion is the Lewis base in this reaction, donating one of its lone pairs as demonstrated in Figure 16.1f:

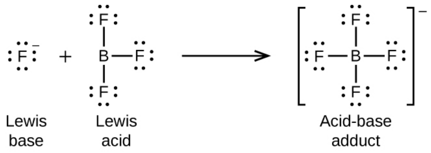


**Figure 16.1f** Lewis Acid-Base Reaction Involving Boron Trifluoride: The Lewis base, a fluorine ion, reacts with a Lewis acid, boron trifluoride, to produce boron tetrafluoride, which is an acid-base adduct. (credit: *Chemistry (OpenStax)*, CC BY 4.0).

Nonmetal oxides act as Lewis acids and react with oxide ions, Lewis bases, to form oxyanions as shown in Figure 16.1g.

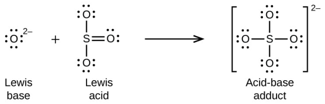


**Figure 16.1g** Lewis Acid-Base Reaction Involving Nonmetal Oxides: The Lewis base, an oxygen ion, reacts with a Lewis acid, sulfur trioxide, to produce a sulphate ion, which is an acid-base adduct. (credit: *Chemistry (OpenStax)*, CC BY 4.0)

Many Lewis acid-base reactions are displacement reactions in which one Lewis base displaces another Lewis base from an acid-base adduct, or in which one Lewis acid displaces another Lewis acid (Figure 16.1h):

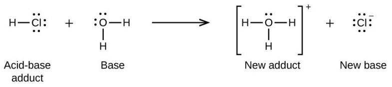


**Figure 16.1h** Lewis Acid-Base Displacement Reaction: In this reaction, HCl is an acid-base adduct that reactants with water, which is behaving as a base, to form a new adduct (hydronium ion) and a new base (chloride ion) (credit: *Chemistry (OpenStax)*, CC BY 4.0)

The displacement reaction in Figure 16.1h shows how the reaction of a Brønsted-Lowry acid with a base fits into the Lewis concept. A Brønsted-Lowry acid such as HCl is an acid-base adduct according to the Lewis concept, and proton transfer occurs because a more stable acid-base adduct is formed. Thus, although the definitions of acids and bases in the two theories are quite different, the theories overlap considerably.

## Amphiprotic / Amphoteric Species

As illustrated in the previous paragraphs water can behave as an acid or a base. Like water, many molecules and ions may either gain or lose a proton under the appropriate conditions. Such species are said to be **amphiprotic**. Another term used to describe such species is **amphoteric**, which is a more general term for a species that may act either as an acid or a base by any definition (not just the Brønsted-Lowry one). Consider for example the bicarbonate ion, which may either donate or accept a proton as shown here:

$$\mathrm{HCO_3}^-(aq) + \mathrm{H_2O}(l) \leftrightharpoons \mathrm{CO_3}^{2-}(aq) + \mathrm{H_3O}^+(aq)$$

$$\mathrm{HCO_3}^-(aq) + \mathrm{H_2O}(l) \leftrightharpoons \mathrm{H_2CO_3}(aq) + \mathrm{OH}^-(aq)$$

### Example 16.1c

### Representing the Acid-Base Behaviour of an Amphoteric Substance

Write separate equations representing the reaction of $HSO_3^-$.

a. as an acid with $OH^-$
b. as a base with HI

### Solution

a. $\mathrm{HSO_3}^-(aq) + \mathrm{OH}^-(aq) \leftrightharpoons \mathrm{SO_3}^{2-}(aq) + \mathrm{H_2O}(l)$
b. $\mathrm{HSO_3}^-(aq) + \mathrm{HI}(aq) \leftrightharpoons \mathrm{H_2SO_3}(aq) + \mathrm{I}^-(aq)$

## Exercise 16.1e

Write separate equations representing the reaction of $H_2PO_4$.

a. as a base with HBr
b. as an acid with $OH^-$

**Check Your Answer**[5]

## Links to Interactive Learning Tools

Explore Acid-Base Properties (//www.physicsclassroom.com/Concept-Builders/Chemistry/Acid-Base-Properties) from the Physics Classroom (https://www.physicsclassroom.com/).

Explore Bronsted-Lowry Acid and Base Model (https://www.physicsclassroom.com/Concept-Builders/Chemistry/Acids-and-Bases) from the Physics Classroom (https://www.physicsclassroom.com/).

## Attribution & References

Except where otherwise noted, this section is adapted by Jackie MacDonald from “14.2: Acids- Properties and Examples“, “14.3: Bases- Properties and Examples“, and “14.7: Strong and Weak Acids and Bases” In *Map: Introductory Chemistry (Tro)* by Marisa Alviar-Agnew & Henry Agnew, Shared under CK-12 license. / Content streamlined and remixed for student comprehension.

- “Molecular Definitions of Acids and Bases” and “Amphiprotic / Amphoteric Species” sections are adapted from “14.1 Brønsted-Lowry Acids and Bases” and “15.2 Lewis Acids and Bases” In *General Chemistry 1 & 2* by Rice University, licensed under CC BY 4.0 International License, Except where otherwise noted. / Wording modifications by Jackie MacDonald.

# Notes

1. Question 1:
    1. base;
    2. base;
    3. acid;
    4. base;
    5. acid;
    6. acid;
    7. acid;
    8. acid;
    9. base;
    10. acid;
    11. base;
    12. acid;

    Question 2: (4) when metals are placed in a basic solution, they react, producing hydrogen bubbles; Question 3: acids are found in citrus fruits and provide the bitter taste; Question 4: FALSE
2. 
    1. when $NH_2^-$ is the acid; $NH_3$ is its conjugate base;
    2. when HBr the acid; $Br^-$ is its conjugate base;
    3. When $HSO_4^-$ is the acid; $SO_4^{2-}$ is the base.
3. 
    a. $H_2O$ is the Brønsted-Lowry acid and $OH^-$ is its conjugate base; $HS^-$ is the Brønsted-Lowry base and $H_2S$ is its conjugate acid;

    b. $H_2PO_4^-$ is the Brønsted-Lowry acid and $HPO_4^{2-}$ is its conjugate base; $OH^-$ is the Brønsted-Lowry base and $H_2O$ is its conjugate acid;

    c. HCl is the Brønsted-Lowry acid and $Cl^-$ is its conjugate base; $H_2PO_4^-$ is the Brønsted-Lowry base and $H_3PO_4$ is its conjugate acid;

    d. $[Al(H_2O)_6]^{3+}$ is the Brønsted-Lowry acid and $[Al(H_2O)_5(OH)]^{2+}$ is its conjugate base; $[Fe(H_2O)_5(OH)]^{2+}$ is the Brønsted-Lowry base and $[Fe(H_2O)_6]^{3+}$ is its conjugate acid.
4. 
    1. (i) Brønsted-Lowry acid = (1) A compound that donates a proton to another compound; (ii) Brønsted-Lowry base = (2) a compound that accepts a proton
    2. $HPO_4^{2-}$
    3. $Cl^-$
    4. 2. Brønsted-Lowry acid $CH_3OH$, conjugate base $CH_3O^-$
5. 
    a. $H_2PO_4{}^-(aq) + HBr(aq) \leftrightharpoons H_3PO_4(aq) + Br^-(aq)$;

    b. $H_2PO_4{}^-(aq) + OH^-(aq) \leftrightharpoons HPO_4{}^{2-}(aq) + H_2O(l)$

# 16.2 REACTIONS OF ACIDS AND BASES

## Learning Objectives

By the end of this section, you will be able to:

- Summarize general reactions of acids and bases

In this section, the more common types of acid-base reactions that take place in aqueous solutions will be summarized. In continuation with the context learned earlier in this chapter, a strong acid is a substance that will dissolve in water to yield hydronium ions, $H_3O^+$. Whereas, a strong base is a substance that will dissolve in water to yield hydroxide ions, $OH^-$. We will learn other definitions of acids and bases in the following sections. The most common bases are ionic compounds composed of alkali or alkaline earth metal cations (groups 1 and 2) combined with the hydroxide ion. For example, NaOH and $Ca(OH)_2$. When these compounds dissolve in water, hydroxide ions are released directly into the solution. These ions are responsible for the characteristic behaviour of acids and bases in aqueous solution.

## Neutralization Reactions

The reaction between an acid and a base is called an acid-base reaction or a **neutralization reaction**. Later in this chapter, a section is dedicated to discussing neutralization reactions and introducing calculations relating to neutralization reactions. However, a brief introduction into this type of reaction is provided. In reactions where the acid is a hydrogen-ion-containing compound and the base is a hydroxide-ion-containing compound, the two react chemically to form water and a salt. The general form for a neutralization reaction is

$$\text{acid} + \text{base} \rightarrow \text{water} + \text{salt}$$

where the term salt is used to define any ionic compound (soluble or insoluble) that is formed from a reaction between an acid and a base, as outlined in chapter 15. In chemistry, the word salt refers to more than just table salt. For example, the balanced chemical equation for the reaction between $HCl(aq)$ and $KOH(aq)$ is

$$HCl(aq) + KOH(aq) \rightarrow H_2O(l) + KCl(aq)$$

where in this instance, the salt is soluble potassium chloride, KCl. By counting the number of atoms of each element, we find that only one water molecule is formed as a product. However, in the reaction between HCl(*aq*) and $Mg(OH)_2$(*aq*), additional molecules of HCl and $H_2O$ are required to balance the chemical equation:

$$2HCl(aq) + Mg(OH)_2(aq) \rightarrow 2H_2O(l) + MgCl_2(aq)$$

Here, the soluble salt is $MgCl_2$. (This is one of several reactions that take place when a type of antacid—a base—is used to treat stomach acid.)

There are acid-base reactions that do not follow the "general acid-base" equation given above. For example, the balanced chemical equation for the reaction between HCl(*aq*) and $NH_3$(*aq*) is

$$HCl(aq) + NH_3(aq) \rightarrow NH_4Cl(aq)$$

Neutralization reactions are one type of chemical reaction that proceeds even if one reactant is not in the aqueous phase. For example, the chemical reaction between HCl(*aq*) and $Fe(OH)_3$(*s*) still proceeds according to the equation

$$3\,HCl(aq) + Fe(OH)_3(s) \rightarrow 3H_2O(l) + FeCl_3(aq)$$

even though $Fe(OH)_3$ is not soluble. When one realizes that $Fe(OH)_3$(*s*) is a component of rust, this explains why some cleaning solutions for rust stains contain acids — the neutralization reaction produces products that are soluble and wash away. Washing with acids like HCl is one way to remove rust and rust stains, but HCl must be used with caution!

## Gas-Forming Acid-Base Reactions

A driving force for certain acid-base reactions is the formation of a gas. Common gases formed are hydrogen ($H_2$), oxygen ($O_2$), and carbon dioxide ($CO_2$).

### Acid Reactions with Carbonates

Many acids react with carbonates to produce a salt, carbon dioxide gas and water.

$$\text{acid} + \text{carbonate} \rightarrow \text{salt} + CO_2(g) + H_2O(l)$$

Consider the following reaction between hydrochloric acid and sodium carbonate (which is considered a base):

$$2HCl(aq) + Na_2CO_3(aq) \rightarrow H_2CO_3(aq) + 2NaCl(aq) \rightarrow CO_2(g) + H_2O(l) + 2NaCl(aq)$$

The above example can be viewed as an acid-base reaction followed by a decomposition. The driving force in this case is the gas formation. The decomposition of $H_2CO_3$ into $CO_2$ and $H_2O$ is a very common reaction. Both sodium carbonate ($Na_2CO_3$) and sodium bicarbonate ($NaHCO_3$) mixed with acid result in a gas-forming acid-base reaction.

$$HCl(aq) + NaHCO_3(aq) \rightarrow H_2CO_3(aq) + NaCl(aq) \rightarrow CO_2(g) + H_2O(l) + NaCl(aq)$$

## More Reactions Involving Acids and Bases

There are many common reactions of acids chemically combining with carbonate, to form a salt and carbonic acid as demonstrated above. Due to the chemical disposition of carbonic acid as weak acid, it readily decomposes to water and carbon dioxide gas. The final products of chemical reactions are shown for examples 1 through 3 below, but remember that the formation of carbonic acid occurs as an intermediate step before decomposing to the final products.

- Example 1: Nitric acid reacts with sodium carbonate to form sodium nitrate, carbon dioxide, and water.

$$2HNO_3(aq) + Na_2CO_3(aq) \rightarrow 2NaNO_3(aq) + CO_2(g) + H_2O(l)$$

- Example 2: Hydrochloric acid reacts with calcium carbonate to form calcium chloride, carbon dioxide, and water:

$$2HCl(aq) + CaCO_3(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

- Example 3: Sulfuric acid reacts with calcium carbonate to form calcium sulfate, carbon dioxide, and water:

$$H_2SO_4(aq) + CaCO_3(aq) \rightarrow CaSO_4(aq) + CO_2(g) + H_2O(l)$$

**Watch Acid + Metal Carbonate | Acids, Bases & Alkalis | Chemistry | FuseSchool (3min 0s) (https://youtu.be/5ynCflOY-LI)**.

### Acid Reactions with Metals

The oxidation of metals in acidic solutions is another chemical process that generates a gas. This reaction will yield a metal salt and hydrogen gas and has the general formula

acid + metal → hydrogen + salt

Consider the following single displacement reaction between hydrochloric acid and solid zinc metal.

$$2HCl(aq) + Zn(s) \rightarrow ZnCl_2(aq) + H_2(g)$$

Here, hydrochloric acid reacts with zinc to produce an aqueous metal salt and hydrogen gas bubbles.

## Base Reactions with Metals

Some bases can also react with metals to form a gaseous product. For instance, strong bases, such as sodium hydroxide and potassium hydroxide, react with amphoteric metals to produce hydrogen gas and a salt:

Base + metal + water → hydrogen + salt

Consider the following example of sodium hydroxide reacting with Zinc in water to form sodium aluminate and hydrogen gas.

$$2NaOH(aq) + Zn(s) \rightarrow Na_2Zn(OH)_4(aq) + H_2(g)$$

## Reactions with Metal Oxides

The oxides of metals are basic and when a metal oxide reacts with an acid, it forms a salt with water.

$$\text{acid} + \text{metal oxide} \rightarrow \text{salt} + H_2O(l)$$

An example of this type of reaction is mixing a small amount of the black, insoluble salt, copper(II) oxide, with dilute sulfuric acid. Upon mixing of the reactants, the colour of the solution becomes blue and the copper oxide dissolves. The blue colour of the solution indicates the formation of the soluble salt, copper(II) sulfate and water.

This chemical reaction can be written as the following:

$$CuO(s) + H_2SO_4\,(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$

**Figure 16.2a** Crystal of Copper(II) Sulfate: Sample of the collection of artificial minerals from the Natural History Museum Lille. This blue crystal of copper(II) sulfate is of industrial origin, created by slow crystallization from a saturated solution. (credit: work by Lamiot, CC BY-SA 4.0).

**Source:** Except where otherwise noted, "More Reactions Involving Acids & Bases" is created by Jackie MacDonald, CC BY-NC-SA 4.0

## Food and Drink: Acids in Foods

Many foods and beverages contain acids. Acids impart a sour note to the taste of foods, which may add some pleasantness to the food. For example, orange juice contains citric acid, $H_3C_6H_5O_7$. Note how this formula shows hydrogen atoms in two places; the first hydrogen atoms written are the hydrogen atoms that can form $H^+$ ions, while the second hydrogen atoms written are part of the citrate ion, $C_6H_5O_7^{3-}$. Lemons and limes contain much more citric acid—about 60 times as much—which accounts for these citrus fruits being more sour than most oranges. Vinegar is essentially a ~5% solution of acetic acid ($HC_2H_3O_2$) in water. Apples contain malic acid ($H_2C_4H_4O_5$; the name *malic acid* comes from the apple's botanical genus name, *malus*), while lactic acid ($HC_3H_5O_3$) is found in wine and sour milk products, such as yogurt and some cottage cheeses.

Table 16.2a "Various Acids Found in Food and Beverages" lists some acids found in foods, either naturally or as an additive. Frequently, the salts of acid anions are used as additives, such as monosodium glutamate (MSG), which is the sodium salt derived from glutamic acid. As you read the list, you should come to the inescapable conclusion that it is impossible to avoid acids in food and beverages.

**Table 16.2a Various Acids Found in Food and Beverages**

| Acid Name | Acid Formula | Use and Appearance |
|---|---|---|
| acetic acid | $HC_2H_3O_2$ | flavouring; found in vinegar |
| adipic acid | $H_2C_6H_8O_4$ | flavouring; found in processed foods and some antacids |
| alginic acid | various | thickener; found in drinks, ice cream, and weight loss products |
| ascorbic acid | $HC_6H_7O_6$ | antioxidant, also known as vitamin C; found in fruits and vegetables |
| benzoic acid | $HC_6H_5CO_2$ | preservative; found in processed foods |
| citric acid | $H_3C_6H_5O_7$ | flavouring; found in citrus fruits |
| dehydroacetic acid | $HC_8H_7O_4$ | preservative, especially for strawberries and squash |
| erythrobic acid | $HC_6H_7O_6$ | antioxidant; found in processed foods |
| fatty acids | various | thickener and emulsifier; found in processed foods |
| fumaric acid | $H_2C_4H_2O_4$ | flavouring; acid reactant in some baking powders |
| glutamic acid | $H_2C_5H_7NO_4$ | flavouring; found in processed foods and in tomatoes, some cheeses, and soy products |
| lactic acid | $HC_3H_5O_3$ | flavouring; found in wine, yogurt, cottage cheese, and other sour milk products |
| malic acid | $H_2C_4H_4O_5$ | flavouring; found in apples and unripe fruit |
| phosphoric acid | $H_3PO_4$ | flavouring; found in some colas |
| propionic acid | $HC_3H_5O_2$ | preservative; found in baked goods |
| sorbic acid | $HC_6H_7O_2$ | preservative; found in processed foods |
| stearic acid | $HC_{18}H_{35}O_2$ | anticaking agent; found in hard candies |
| succinic acid | $H_2C_4H_4O_4$ | flavouring; found in wine and beer |
| tartaric acid | $H_2C_4H_4O_6$ | flavouring; found in grapes, bananas, and tamarinds |

## Attribution & References

Except where otherwise noted, this page was adapted by Jackie MacDonald from from “4.3 Acid-Base Reactions” In *Introduction to Chemistry* by Carol Higginbotham, licensed under CC BY-NC-SA. / Sections “Acid Reactions with Metals”, “Base Reactions with Metals” and “Reactions with Metal Oxides” written by Jackie MacDonald.

# 16.3 IONIZATION OF WATER

## Learning Objectives

By the end of this section, you will be able to:

- Describe the autoionization of water
- Use the ion-product constant for water to calculate hydronium and hydroxide ion concentrations

To begin this chapter, **Watch How Polarity makes Water Behave Strangely (3min 51s). (https://youtu.be/ASLUY2U1M-8)**

Earlier in this chapter, the concept of amphoteric species, such as water, was introduced. Because of its highly polar structures, water can function as either an acid or a base, depending on the nature of the solute dissolved in it. In fact, in pure water or in any aqueous solution, water acts both as an acid and a base. A very small fraction of water molecules donate protons to other water molecules to form hydronium ions and hydroxide ions (Figure 16.3a).

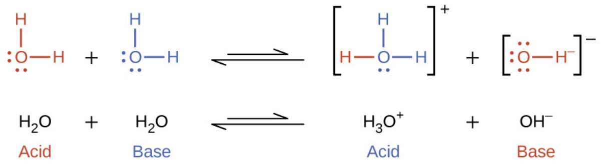


**Figure 16.3a** Autoionization of Water: Water acts both as an acid and a base. A very small fraction of water molecules donate protons to other water molecules to form hydronium ions and hydroxide ions (credit: *Chemistry (OpenStax)*, CC BY 4.0).

This type of reaction, in which a substance ionizes when one molecule of the substance reacts with another molecule of the same substance, is referred to as **autoionization**.

To visualize this process in a video simulation, **watch Autoionization in Liquid Water (3 min 23s). (https://youtu.be/zeFSzt5x9uo)**

Pure water undergoes autoionization to a very slight extent. Only about two out of every $10^9$ molecules in a sample of pure water are ionized at 25 °C. The relationship between products and reactants of any reaction at equilibrium can be expressed by its equilibrium constant, K. The equilibrium constant for the ionization of water is called the **ion-product constant for water ($K_w$)**:

$$\mathrm{H_2O}(l) \;+\; \mathrm{H_2O}(l) \rightleftharpoons \mathrm{H_3O^+}(aq) \;+\; \mathrm{OH^-}(aq)$$

$$K_\mathrm{w} = [\mathrm{H_3O^+}][\mathrm{OH^-}]$$

The slight ionization of pure water is reflected in the small value of the equilibrium constant; at 25 °C, $K_w$ has a value of $1.0 \times 10^{-14}$. The process is endothermic, and so the extent of ionization and the resulting concentrations of hydronium ion and hydroxide ion increase with temperature. For example, at 100 °C, the value for $K_w$ is about $5.6 \times 10^{-13}$, roughly 50 times larger than the value at 25 °C.

## Example 16.3a

### Ion Concentrations in Pure Water

What are the hydronium ion concentration and the hydroxide ion concentration in pure water at 25 °C?

### Solution

The autoionization of water yields the same number of hydronium and hydroxide ions. Therefore, in pure water, $[H_3O^+] = [OH^-]$. At 25 °C:

$$K_\mathrm{w} = [\mathrm{H_3O^+}][\mathrm{OH^-}] = [\mathrm{H_3O^+}]^2 = [\mathrm{OH^-}]^2 = 1.0 \;\times\; 10^{-14}$$

So:

$$[\mathrm{H_3O^+}] = [\mathrm{OH^-}] = \sqrt{1.0 \;\times\; 10^{-14}} = 1.0 \;\times\; 10^{-7}\; M$$

The hydronium ion concentration and the hydroxide ion concentration are the same, and we find that both equal $1.0 \times 10^{-7}$ *M*.

Therefore, at 25 °C, $K_w$ has a value of $1.0 \times 10^{-14}$.

$$K_\mathrm{w} = [\mathrm{H_3O^+}][\mathrm{OH^-}] = 1.0 \;\times\; 10^{-14}$$

The degree of autoionization of water and hence the value of $Kw$ changes with temperature, so the equation

$$K_\mathrm{w} = [\mathrm{H_3O^+}][\mathrm{OH^-}] = 1.0 \;\times\; 10^{-14}$$

is accurate only at room temperature. If a temperature is not given in a question, assume room temperature.

### Exercise 16.3a

The ion product of water at 80 °C is $2.4 \times 10^{-13}$. What are the concentrations of hydronium and hydroxide ions in pure water at 80 °C?

**Check Your Answer**[1]

## Calculating Hydronium and Hydroxide Concentrations in Solutions using $K_w$

It is important to realize that the autoionization equilibrium for water is established in all aqueous solutions. Pure water is considered a **neutral solution**. Adding an acid or base to water will not change the position of the equilibrium.

### Ion Concentrations in Water

At 25°C,

- In a neutral solution, the $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ $M$
- In an acidic solution, the concentration of $H_3O^+(aq)$ is greater than $1.0 \times 10^{-7}$ $M$, and thus, its $[OH^-]$ will be less than $1.0 \times 10^{-7}$ $M$
    - In an acidic solution, the $[H_3O^+] > [OH^-]$
- For basic solutions, the concentration of $OH^-(aq)$ is greater than $1.0 \times 10^{-7}$ $M$, and thus, its $[H_3O^+]$ will be less than $1.0 \times 10^{-7}$ $M$
    - In a basic solution, the $[OH^-] > [H_3O^+]$

However, remember the product of the two concentrations $[H_3O^+][OH^-]$ is always equal to $1.0 \times 10^{-14}$ $M$, no matter whether the aqueous solution is an acid, a base, or neutral:

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}\ M$$

Consequently, if you know $[H_3O^+]$ for a solution, you can use the $K_w$ formula to calculate the $[OH^-]$.

Alternatively, if you know [$OH^-$], you can calculate [$H_3O^+$]. This means in any given acid/base solution the two concentrations are inversely related: As one concentration increases, the other must decrease, so their product always equals the value of $K_w$.

**Source:** "Ion Concentrations in Water" by Jackie MacDonald, CC BY-NC-SA 4.0

Example 16.3b demonstrates the quantitative aspects of this relation between hydronium and hydroxide ion concentrations. At 25°C, $K_{\mathrm{w}} = [\mathrm{H_3O^+}][\mathrm{OH^-}] = 1.0 \times 10^{-14}$.

## Example 16.3b

### The Inverse Proportionality of [$H_3O^+$] and [$OH^-$]

A solution of carbon dioxide in water has a hydronium ion concentration of 2.0 × $10^{-6}$ *M*. What is the concentration of hydroxide ion at 25 °C?

### Solution

We know the value of the ion-product constant for water at 25 °C:

$$2\mathrm{H_2O}(l) \leftrightharpoons \mathrm{H_3O^+}(aq) + \mathrm{OH^-}(aq)$$

$$K_{\mathrm{w}} = [\mathrm{H_3O^+}][\mathrm{OH^-}] = 1.0 \times 10^{-14}$$

Thus, we can calculate the missing equilibrium concentration.

Rearrangement of the $K_w$ expression yields that [$OH^-$] is directly proportional to the inverse of [$H_3O^+$]:

$$[\mathrm{OH^-}] = \frac{K_{\mathrm{w}}}{[\mathrm{H_3O^+}]} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-6}} = 5.0 \times 10^{-9}$$

The hydroxide ion concentration in water is reduced to 5.0 × $10^{-9}$ *M* as the hydrogen ion concentration increases to 2.0 × $10^{-6}$ *M*. This is expected from Le Châtelier's principle; the autoionization reaction shifts to the left to reduce the stress of the increased hydronium ion concentration and the [$OH^-$] is reduced relative to that in pure water.

A check of these concentrations confirms that our arithmetic is correct:

$$K_{\mathrm{w}} = [\mathrm{H_3O^+}][\mathrm{OH^-}] = (2.0 \times 10^{-6})(5.0 \times 10^{-9}) = 1.0 \times 10^{-14}$$

## Exercise 16.3b

What is the hydronium ion concentration in an aqueous solution with a hydroxide ion concentration of 0.001 *M* at 25 °C?

**Check Your Answer**[2]

## Exercise 16.3c

**Practice using the following PhET simulation: pH Scale (https://phet.colorado.edu/sims/html/ph-scale/latest/ph-scale_en.html)**

## Example 16.3c

For the following solutions, determine whether the solution is acidic, basic or neutral.

1. A solution that has a $[H_3O^+] = 3.5 \times 10^{-3}\ M$
2. A solution that has a $[OH^-] = 1.0 \times 10^{-7}\ M$
3. A solution that has a $[H_3O^+] = 3.5 \times 10^{-12}\ M$

### Solution

1. Since the solution's $[H_3O^+]$ of $3.5 \times 10^{-3}\ M$ is greater than $1.0 \times 10^{-7} M$, the solution is acidic.
2. Since the solution's $[OH^-]$ of $1.0 \times 10^{-7} M$ is equal to $1.0 \times 10^{-7} M$, the solution is neutral.
3. Since the solution's $[H_3O^+]$ of $3.5 \times 10^{-12}\ M$ is less than $1.0 \times 10^{-7} M$, its $[OH^-]$ will be greater $1.0 \times 10^{-7} M$, which indicates the solution is basic.

**Activity Source:** "Example 16.3c" created by Jackie MacDonald is licensed under CC BY-NC-SA 4.0.

## Exercise 16.3d

Check Your Learning Exercise (Text Version)
For the following solutions, determine whether the solution is acidic, basic or neutral.

1. A solution that has a $[H_3O^+] = 2.5 \times 10^{-12}\ M$
2. A solution that has a $[H_3O^+] = 9.0 \times 10^{-4}\ M$
3. A solution that has a $[H_3O^+] = 1.0 \times 10^{-7}\ M$
4. A solution that has a $[OH^-] = 1.0 \times 10^{-9}\ M$
5. A solution that has a $[OH^-] = 1.0 \times 10^{-7}\ M$
6. A solution that has a $[OH^-] = 4.8 \times 10^{-2}\ M$

**Check Your Answer**[3]

**Source:** "Exercise 16.3d" by Jackie MacDonald, licensed under CC BY NC SA 4.0.

## Key Equations

- $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ (at 25 °C)

## Attribution & References

Except where otherwise noted, this page is adapted from "14.1 Brønsted-Lowry Acids and Bases" In *General Chemistry 1 & 2* by Rice University, a derivative of *Chemistry (Open Stax)* by Paul Flowers, Klaus Theopold, Richard Langley & William R. Robinson and is licensed under CC BY 4.0. Access for free at *Chemistry (OpenStax) (https://openstax.org/books/chemistry/pages/1-introduction)*

## Notes

1. $[H_3O^+] = [OH^-] = 4.9 \times 10^{-7} M$
2. $[H_3O^+] = 1 \times 10^{-11} M$
3. (1) Basic; (2) Acidic; (3) Neutral; (4) Acidic; (5) Neutral; (6) Basic

# 16.4 INTRODUCTION TO PH AND POH

## Learning Objectives

By the end of this section, you will be able to:

- Explain the characterization of aqueous solutions as acidic, basic, or neutral
- Express hydronium and hydroxide ion concentrations on the pH and pOH scales
- Perform calculations relating pH and pOH

As discussed earlier, hydronium and hydroxide ions are present both in pure water and in all aqueous solutions, and their concentrations are inversely proportional as determined by the ion product of water ($K_w$). The concentrations of these ions in a solution are often critical determinants of the solution's properties and the chemical behaviours of its other solutes, and specific vocabulary has been developed to describe these concentrations in relative terms. A solution is **neutral** if it contains equal concentrations of hydronium and hydroxide ions; **acidic** if it contains a greater concentration of hydronium ions than hydroxide ions; and **basic** if it contains a lesser concentration of hydronium ions than hydroxide ions. To assist with informing users of a substance's acidity or basicity (alkalinity), one can calculate the pH of the substance and reference it on the pH scale. For a brief introduction into the pH scale and how it is used, **Watch The pH Scale Explained (5min 54s) (https://youtu.be/Xkrgw2CFwRs).**

## Introduction to the pH Scale

The **pH scale** is a scale of acidity ranging 0 to 14 (but not always). It indicates how acidic or basic a solution is. You can use **pH** to make a quick determination whether a given aqueous solution is acidic, basic, or neutral. Since the pH of a substance is dependent on hydronium ion concentration, $[H_3O^+]$, the following conclusions (at standard temperature of 25 °C) can be inferred:

- If pH $< 7$, then the solution is acidic.

- If pH = 7, then the solution is neutral.
- If pH > 7, then the solution is basic.

Most substances have a pH in the range of 0 to 14, although *extremely* acidic or alkaline substances may have pH < 0, or pH > 14, respectively.

**Source:** "pH" by Simple English Wikipedia, licensed under CC BY-SA 3.0.

### Exercise 16.4a

Check Your Learning Exercise (Text Version)
Label each solution as acidic, basic, or neutral based only on the given pH value:

1. Household bleach, pH = 12.6
2. Lime Juice, pH = 2.3
3. Milk of Magnesia, pH = 10.5
4. Pure Water, pH = 7
5. Vinegar, pH = 2.5
6. Baking soda, pH = 8.5

**Check Your Answer**[1]

**Source:** "Exercise 16.4a" by Jackie MacDonald, licensed under CC BY-NC-SA 4.0.

## Formulas for Calculating pH, pOH, [$H_3O^+$] and [$OH^-$]

A common means of expressing quantities, the values of which may span many orders of magnitude, is to use a logarithmic scale. One such scale that is very popular for chemical concentrations and equilibrium constants is based on the p-function, defined as shown where "X" is the quantity of interest and "log" is the base-10 logarithm:

$$\mathrm{pX} = -\log \mathrm{X}$$

The pH of a solution is therefore defined as shown here, where [$H_3O^+$] is the molar concentration of hydronium ion in the solution:

$$\mathrm{pH} = -\log[\mathrm{H_3O^+}]$$

Rearranging this equation to isolate the hydronium ion molarity yields the equivalent expression:

$$[\mathrm{H_3O^+}] = 10^{-\mathrm{pH}}$$

Likewise, the hydroxide ion molarity may be expressed as a p-function, or **pOH**:

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

or

$$[\text{OH}^-] = 10^{-\text{pOH}}$$

Finally, the relation between these two ion concentration expressed as p-functions is easily derived from the $K_w$ expression:

$$K_w = [H_3O^+][OH^-]$$
$$-\log K_w = -\log([H_3O^+][OH^-]) = -\log[H_3O^+] + -\log[OH^-]$$
$$pK_w = pH + pOH$$

At 25 °C, the value of $K_w$ is $1.0 \times 10^{-14}$, and so:

$$14.00 = pH + pOH$$

As was shown in Example 16.3a in Chapter 16.3 Ionization of Water, Brønsted-Lowry Acids and Bases, the hydronium ion molarity in pure water (or any neutral solution) is $1.0 \times 10^{-7} M$ at 25 °C. The pH and pOH of a neutral solution at this temperature are therefore:

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(1.0 \times 10^{-7}) = 7.00$$
$$\text{pOH} = -\log[\text{OH}^-] = -\log(1.0 \times 10^{-7}) = 7.00$$

And so, *at this temperature*, acidic solutions are those with hydronium ion molarities greater than $1.0 \times 10^{-7} M$ and hydroxide ion molarities less than $1.0 \times 10^{-7} M$ (corresponding to pH values less than 7.00 and pOH values greater than 7.00). Basic solutions are those with hydronium ion molarities less than $1.0 \times 10^{-7} M$ and hydroxide ion molarities greater than $1.0 \times 10^{-7} M$ (corresponding to pH values greater than 7.00 and pOH values less than 7.00).

Since the autoionization constant $K_w$ is temperature dependent, these correlations between pH values and the acidic/neutral/basic adjectives will be different at temperatures other than 25 °C. For example, the hydronium molarity of pure water at 80 °C is $4.9 \times 10^{-7} M$, which corresponds to pH and pOH values of:

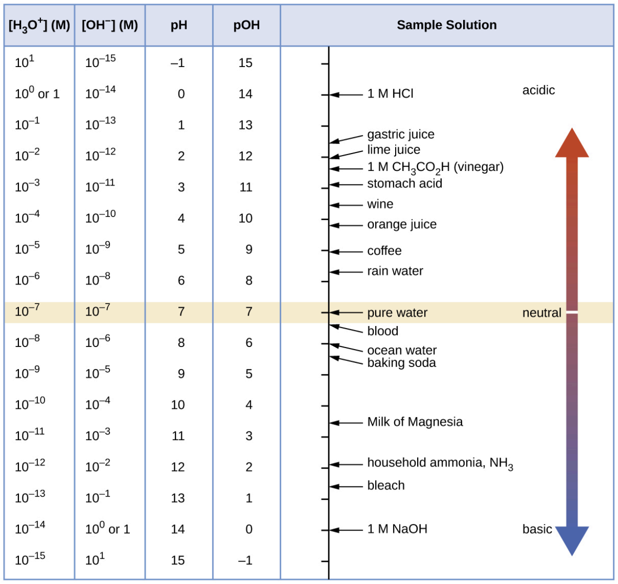
$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(4.9 \times 10^{-7}) = 6.31$$
$$\text{pOH} = -\log[\text{OH}^-] = -\log(4.9 \times 10^{-7}) = 6.31$$

At this temperature, then, neutral solutions exhibit pH = pOH = 6.31, acidic solutions exhibit pH less than 6.31 and pOH greater than 6.31, whereas basic solutions exhibit pH greater than 6.31 and pOH less than 6.31. This distinction can be important when studying certain processes that occur at nonstandard temperatures, such as enzyme reactions in warm-blooded organisms. Unless otherwise noted, references to pH values are presumed to be those at standard temperature (25 °C) (Table 16.4a).

**Table 16.4a Summarizing pH Properties for Acidic, Basic and Neutral Solutions**

| Classification | Relative Ion Concentrations | pH at 25 °C |
|---|---|---|
| acidic | $[H_3O^+] > [OH^-]$ | pH < 7 |
| neutral | $[H_3O^+] = [OH^-]$ | pH = 7 |
| basic | $[H_3O^+] < [OH^-]$ | pH > 7 |

Figure 16.4a shows the relationships between $[H_3O^+]$, $[OH^-]$, pH, and pOH, and gives values for these properties at standard temperatures for some common substances.



| $[H_3O^+]$ (M) | $[OH^-]$ (M) | pH | pOH |
|---|---|---|---|
| $10^1$ | $10^{-15}$ | –1 | 15 |
| $10^0$ or 1 | $10^{-14}$ | 0 | 14 |
| $10^{-1}$ | $10^{-13}$ | 1 | 13 |
| $10^{-2}$ | $10^{-12}$ | 2 | 12 |
| $10^{-3}$ | $10^{-11}$ | 3 | 11 |
| $10^{-4}$ | $10^{-10}$ | 4 | 10 |
| $10^{-5}$ | $10^{-9}$ | 5 | 9 |
| $10^{-6}$ | $10^{-8}$ | 6 | 8 |
| $10^{-7}$ | $10^{-7}$ | 7 | 7 |
| $10^{-8}$ | $10^{-6}$ | 8 | 6 |
| $10^{-9}$ | $10^{-5}$ | 9 | 5 |
| $10^{-10}$ | $10^{-4}$ | 10 | 4 |
| $10^{-11}$ | $10^{-3}$ | 11 | 3 |
| $10^{-12}$ | $10^{-2}$ | 12 | 2 |
| $10^{-13}$ | $10^{-1}$ | 13 | 1 |
| $10^{-14}$ | $10^0$ or 1 | 14 | 0 |
| $10^{-15}$ | $10^1$ | 15 | –1 |

**Figure 16.4a** Relationships between $[H_3O^+]$, $[OH^-]$, pH, and pOH: The pH and pOH scales represent concentrations of $[H_3O^+]$ and $OH^-$, respectively. The pH and pOH values of some common substances at standard temperature (25 °C) are shown in this chart (credit: *Chemistry (OpenStax)*, CC BY 4.0).

## Determining pH and pOH from its Reciprocal pH/pOH Value

pOH scale is fairly similar to pH, except instead of measuring the hydrogen ion concentration in moles per litre, it measures the hydroxide ion concentration in moles per litre. As shown earlier in this section (during standard temperature conditions), pH and pOH must add up to 14:

$$14.00 = \text{pH} + \text{pOH}$$

So if you are given pH, you can determine pOH by rearranging the formula:

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

Alternatively, if you are given pOH, you can determine pH by rearranging the formula:

$$pH = 14 - pOH$$

## Example 16.4a

A cleaning solution has a pOH of 2.5. Calculate the pH of this solution and determine whether this solution is acidic, basic, or neutral.

**Solution**

Given information

pOH = 2.5

To find the pH, use the formula

pH = 14 – pOH

Solve:

pH = 14 – 2.5

pH = 11.5; this cleaning solution is basic since its pH is greater than 7.

## Exercise 16.4b

### Exercise 16.4b Part 1:

For each solution listed below, calculate its pOH and determine whether this solution is acidic, basic, or neutral.

a. Solution that has a pH = 9.25
b. Solution that has a pH = 3.8

### Exercise 16.4b Part 2:

For each solution listed below, calculate its pH and determine whether this solution is acidic, basic, or neutral.

a. Solution that has a pOH = 5.65
b. Solution that has a pOH = 13.1

### Check Your Answer[2]

Essentially, the scale for pOH is the reverse of the pH scale.

**Source:** "Determining pH and pOH from its Reciprocal pH/pOH Value" by Jackie MacDonald, CC-BY-NC-SA 4.0

The acidity of a solution is typically assessed experimentally by measurement of its pH. The pOH of a solution is not usually measured, as it is easily calculated from an experimentally determined pH value. The pH of a solution can be directly measured using a pH meter (Figure 16.4c).

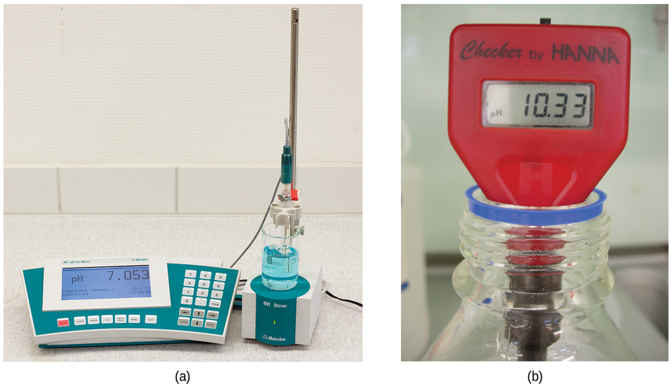


**Figure 16.4b** (a) A research-grade pH meter used in a laboratory can have a resolution of 0.001 pH units, an accuracy of ± 0.002 pH units, and may cost in excess of 1000 dollars. (b) A portable pH meter has lower resolution (0.01 pH units), lower accuracy (± 0.2 pH units), and a far lower price tag. (credit a: work by Datamax, PD; credit b: modification of work by Laurence Livermore (https://www.flickr.com/photos/82012125@N00), CC BY 2.0)

The pH of a solution may also be visually estimated using coloured indicators (Figure 16.4d).

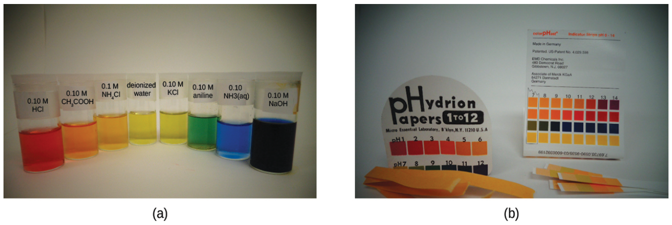


**Figure 16.4c** (a) A universal indicator assumes a different colour in solutions of different pH values. Thus, it can be added to a solution to determine the pH of the solution. The eight vials each contain a universal indicator and 0.1-M solutions of progressively weaker acids: HCl (pH = 1), CH3CO2H (pH = 3), and NH4Cl (pH = 5), deionized water, a neutral substance (pH = 7); and 0.1-M solutions of the progressively stronger bases: KCl (pH = 7), aniline, C6H5NH2 (pH = 9), NH3 (pH = 11), and NaOH (pH = 13). (b) pH paper contains a mixture of indicators that give different colours in solutions of differing pH values. (credit: modification of work by Sahar Atwa in *Chemistry (OpenStax)*, CC BY 4.0).

## Calculating pH when given Hydronium Concentration

When you are given a solution's hydronium concentration, its pH can be calculated using the formula:

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

### Example 16.4b

**Calculation of pH from [H₃O⁺]**

What is the pH of stomach acid, a solution of HCl with a hydronium ion concentration of $1.2 \times 10^{-3}$ *M*?

**Solution**

$$\begin{aligned}\text{pH} &= -\log[\text{H}_3\text{O}^+] \\ &= -\log(1.2 \times 10^{-3}) \\ &= -(-2.92) \\ &= 2.92\end{aligned}$$

(The use of logarithms is explained in Appendix B – The Use of Logarithms and Exponential Numbers Section).

Recall that, as we have done here, when taking the log of a value, keep as many decimal places in the result as there are significant figures in the value.). To review significant figures for logarithm calculations, **Watch Significant Figures and Logarithms (2min 0s). (https://youtu.be/mEwY4f4Tync)**

## Exercise 16.4c

Water exposed to air contains carbonic acid, $H_2CO_3$, due to the reaction between carbon dioxide and water:

$$\mathrm{CO_2}(aq) + \mathrm{H_2O}(l) \leftrightharpoons \mathrm{H_2CO_3}(aq)$$

Air-saturated water has a hydronium ion concentration caused by the dissolved $CO_2$ of $2.0 \times 10^{-6}$ *M*, about 20-times larger than that of pure water. Calculate the pH of the solution at 25 °C.

**Check Your Answer**[3]

## Calculating Hydronium Concentration when given the pH

When you are given a solution's pH, its hydronium concentration can be calculated using the formula:
$[\mathrm{H_3O^+}] = 10^{-\mathrm{pH}}$

### Example 16.4c

**Calculation of Hydronium Ion Concentration from pH**
Calculate the hydronium ion concentration of blood that has a pH of 7.3 (slightly alkaline).

**Solution**

Use the formula:

$$[\mathrm{H_3O^+}] = 10^{-\mathrm{pH}}$$

Solve given that the pH = 7.3

$[H_3O^+] = 10^{-7.3}$

$[H_3O^+] = 5 \times 10^{-8}$ *M*

Therefore, the hydronium ion concentration of blood with a pH of 7.3 is $5 \times 10^{-8}$ *M*

**Source:** "Example 16.4c" by Jackie MacDonald, CC-BY-NC-SA 4.0

### Exercise 16.4d

Calculate the hydronium ion concentration of a solution with a pH of −1.07.

#### Check Your Answer[4]

**Source:** “Exercise 16.4d” by Jackie MacDonald, CC-BY-NC-SA 4.0

## Calculating pOH when given Hydroxide Concentration

When you are given a solution’s hydronium concentration, its pH can be calculated using the formula: $\text{pOH} = -\log[\text{OH}^-]$

### Example 16.4d

What are the pOH and the pH of a 0.0125 *M* solution of potassium hydroxide, KOH?

#### Solution

Potassium hydroxide is a highly soluble ionic compound and completely dissociates when dissolved in dilute solution, yielding $[\text{OH}^-]$ = 0.0125 *M*:

$$\begin{aligned}\text{pOH} &= -\log[\text{OH}^-] \\ &= -\log(0.0125) \\ &= -(-1.903) \\ &= 1.903\end{aligned}$$

The pH can be found from the pOH:

$$\begin{aligned} \text{pH} + \text{pOH} &= 14.00 \\ \text{pH} &= 14.00 - \text{pOH} \\ &= 14.00 - 1.903 \\ &= 12.10 \end{aligned}$$

## Exercise 16.4e

The hydroxide concentration of a sodium hydroxide solution is 0.091 *M*. Determine the pOH and pH of the solution.

**Check Your Answer**[5]

## Calculating Hydroxide Concentration when given the pOH

When you are given a solution's hydronium concentration, its pH can be calculated using the formula: $[\text{OH}^-] = 10^{-\text{pOH}}$

### Example 16.4e

**Calculation of [OH¯]**

The pOH of house hold bleach is 1.45. What is the [OH¯] of this solution?

**Solution**

Use the formula $[\text{OH}^-] = 10^{-\text{pOH}}$ to determine the hydroxide concentration of the bleach solution.

$[\text{OH}^-] = 10^{-\text{pOH}}$

$[\text{OH}^-] = 10^{-1.45}$

$[OH^-]$ = 0.035 *M* or 3.5 x $10^{-2}$ *M*

Source: “Example 16.4e” by Jackie MacDonald, CC-BY-NC-SA 4.0

## Exercise 16.4f

The pOH of toilet a bowl cleaner is measured to be 12.35. What is the $[OH^-]$ of this solution? Is this solution acidic or basic?

**Check Your Answer[6] Source: “Exercise 16.5f” by Jackie MacDonald, CC-BY-NC-SA 4.0**

## Acid/Base Calculations involving Two Steps

Figure 16.4b below shows all of the calculation interrelationships between $[H_3O^+]$, $[OH^-]$, pH, and pOH. The flow chart can be referenced to support students in determining which formula(s) need to be used to solve a given problem.

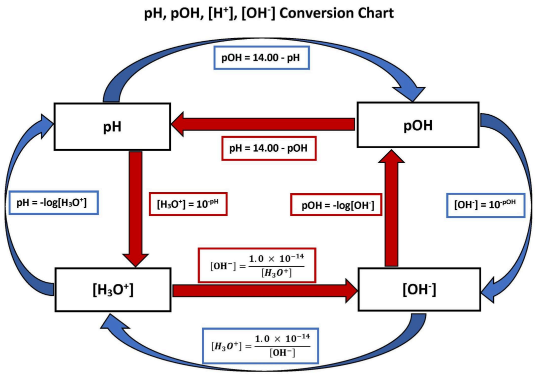


**Figure 16.4d** Calculation Relationships between Hydrogen Ion Concentration, Hydroxide Concentration, pH and pOH: Use the flow chart to determine which formula(s) to use to answer an acid/base calculation. To begin, locate the unit of the given information on the diagram that was provided in the question. Next, locate the unit of the answer you are asked to calculate. Use the formula(s) to solve the question that is associated with the arrow(s) pointing in the direction of the unit you are trying to calculate. It may be a one step or two step problem. (credit: Figure 16.4b by Jackie MacDonald, CC-BY-NC-SA 4.0).

## Example 16.4f

Acidic soils typically yield blue or lavender-blue hydrangea blooms. Alkaline (basic) soil tends to grow pinkish-red blooms. If soil has a $[OH^-]$ of $6.3 \times 10^{-7}$ *M*, what is pH of the soil and what colour blooms will this soil promote for hydrangeas growing in this soil?

### Solution:

Given information: $[OH^-] = 6.3 \times 10^{-7}$ *M*

Asked to find: the pH of soil and, thus, the likely colour of the flower's blooms

Steps – First find pOH, then calculate pH. If acidic, blooms will be blue; if alkaline, blooms will be pink/red.

Step 1 – Calculate pOH using $pOH = -log[OH^-]$

$pOH = -log[6.3 \times 10^{-7}]$

$pOH = 6.20$

Step 2 – Find the pH and determine colour of the blooms

$pH = 14 – pOH$

$pH = 14 – 6.20$

$pH = 7.80$

Therefore the pH of the soil is 7.80, which is slightly alkaline, so the blooms will be a pinkish/red colour.

Source: "Example 16.4e" by Jackie MacDonald, CC BY-NC-SA 4.0

## Exercise 16.4g

The hydronium ion concentration of vinegar is approximately $4 \times 10^{-3}$ *M*. What is the pOH of this solution?

**Check Your Answer**[7] **Source: "Exercise 16.4g" by Jackie MacDonald, CC-BY-NC-SA 4.0**

## Links to Interactive Learning Tools

Explore pH and pOH from the Physics Classroom.

## Attribution & References

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## Notes

1. (1) Basic; (2) Acidic; (3) Basic; (4) Neutral; (5) Acidic; (6) Basic

2. Part 1 (a) pOH = 4.75, basic; (b) pOH = 10.2, acidic; Part 2 (a) pH = 8.35, basic; (b) pH = 0.9, acidic
3. pH = 5.70
4. $[H_3O^+]$ = 12 *M*
5. Step 1: Find pOH using $\mathrm{pOH} = -\log[\mathrm{OH}^-]$; therefore, pOH = 1.04; pH = 14 - 1.04 = 12.96
6. Use the formula $[OH^-] = 10^{-pOH}$ to determine the hydroxide concentration of the bleach solution. $[OH^-] = 10^{-pOH}$ $[OH^-] = 10^{-12.35}$ $[OH^-] = 4.5 \times 10^{-13}$ *M*; since the pOH is higher than 7, this means its corresponding pH is below 7, and the solution is acidic.
7. Step 1: Find pH first using $pH = -\log[H_3O^+]$; pH = 2.4; Step 2 calculate pOH: pOH = 14 - pH = 14 - 2.4 = 11.6

# 16.5 NEUTRALIZATION

## Learning Objectives

By the end of this section, you will be able to:

- Describe a neutralization reaction.
- Predict whether a salt solution will be acidic, basic, or neutral.
- Perform calculations involving titrations of strong acids and strong bases.

As we have seen in the section on chemical reactions, when an acid and base are mixed, they undergo a neutralization reaction. The word "neutralization" seems to imply that a stoichiometrically equivalent solution of an acid and a base would be neutral. This is sometimes true, but the salts that are formed in these reactions may have acidic or basic properties of their own that affect the overall pH.

## Acid-Base Neutralization

Recall the general chemical equation for a neutralization reaction:

$$\text{Acid} \ + \ \text{Base} \rightarrow \text{Salt} \ + \ \text{Water}$$

It is important to remember neutralization reactions are just a specific type of double displacement reaction.

A solution is neutral when it contains equal concentrations of hydronium and hydroxide ions. When we mix solutions of an acid and a base, an acid-base neutralization reaction occurs. However, even if we mix stoichiometrically equivalent quantities, we may find that the resulting solution is not neutral. It could contain either an excess of hydronium ions or an excess of hydroxide ions because the nature of the salt formed determines whether the solution is acidic, neutral, or basic. The following four situations illustrate how solutions with various pH values can arise following a neutralization reaction using stoichiometrically equivalent quantities:

- A strong acid and a strong base, such as $HCl(aq)$ and $NaOH(aq)$, react to form a neutral solution since they fully dissociate in water and their conjugate partners produced are of negligible strength:

$$\mathrm{HCl}(aq)\ +\ \mathrm{NaOH}(aq) \rightleftharpoons \mathrm{NaCl}(aq)\ +\ \mathrm{H_2O}(l)$$

- A strong acid and a weak base yield a weakly acidic solution, not because of the strong acid involved, but because of the conjugate acid of the weak base. A weak base produces a strong conjugate acid, and this will affect the overall pH to be more acidic.
- A weak acid and a strong base yield a weakly basic solution. A solution of a weak acid reacts with a solution of a strong base to form the conjugate base of the weak acid and the conjugate acid of the strong base. The conjugate acid of the strong base is a weaker acid than water and has no effect on the acidity of the resulting solution. However, the conjugate base of the weak acid is a weak base and ionizes slightly in water. This increases the amount of hydroxide ion in the solution produced in the reaction and renders it slightly basic.
- A weak acid plus a weak base can yield either an acidic, basic, or neutral solution. This is the most complex of the four types of reactions. When the conjugate acid and the conjugate base are of unequal strengths, the solution can be either acidic or basic, depending on the relative strengths of the two conjugates. Occasionally the weak acid and the weak base will have the *same* strength, so their respective conjugate base and acid will have the same strength, and the solution will be neutral. To predict whether a particular combination will be acidic, basic or neutral, tabulated $K$ values of the conjugates must be compared.
    - The **acid dissociation constant, (Ka)**, is an equilibrium constant that gives numerical representation of an acid's strength in a solution based its degree of dissociation in water. The greater the $K_a$, the stronger the acid.
    - The **base dissociation constant, ($K_b$)**, is an equilibrium constant that measures how completely a base dissociates into ions in water. The greater the Kb, the stronger the base.

## Stomach Antacids

Our stomachs contain a solution of roughly 0.03 *M* HCl, which helps us digest the food we eat. The burning sensation associated with heartburn is a result of the acid of the stomach leaking through the muscular valve at the top of the stomach into the lower reaches of the esophagus. The lining of the esophagus is not protected from the corrosive effects of stomach acid the way the lining of the stomach is, and the results can be very painful. When we have heartburn, it feels better if we reduce the excess acid in the esophagus by taking an antacid. As you may have guessed, antacids are bases. One of the most common antacids is calcium carbonate, $CaCO_3$. The reaction,

$$\mathrm{CaCO_3}(s)\ +\ 2\mathrm{HCl}(aq) \rightleftharpoons \mathrm{CaCl_2}(aq)\ +\ \mathrm{H_2O}(l)\ +\ \mathrm{CO_2}(g)$$

not only neutralizes stomach acid, it also produces $CO_2(g)$, which may result in a satisfying belch.

Milk of Magnesia is a suspension of the sparingly soluble base magnesium hydroxide, $Mg(OH)_2$. It works according to the reaction:

$$\mathrm{Mg(OH)_2}(s) \rightleftharpoons \mathrm{Mg^{2+}}(aq) \ + \ \mathrm{2OH^-}(aq)$$

The hydroxide ions generated in this equilibrium then go on to react with the hydronium ions from the stomach acid, so that :

$$\mathrm{H_3O^+} \ + \ \mathrm{OH^-} \rightleftharpoons \mathrm{2H_2O}(l)$$

This reaction does not produce carbon dioxide, but magnesium-containing antacids can have a laxative effect.

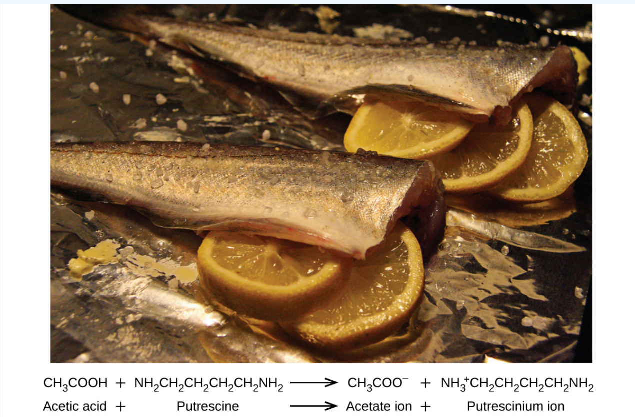
Several antacids have aluminum hydroxide, $Al(OH)_3$, as an active ingredient. The aluminum hydroxide tends to cause constipation, and some antacids use aluminum hydroxide in coordination with magnesium hydroxide to balance the side effects of the two substances.

**Figure 16.5a** Magic of Antacids: A jar of Antacids is shown. Antacids are bases, and are used to neutralize stomach acids by individuals experiencing acid reflux (credit: work by Midnightcomm, CC BY-SA 2.5).

## Culinary Aspects of Chemistry

Cooking is essentially synthetic chemistry that happens to be safe to eat. There are a number of examples of acid-base chemistry in the culinary world. One example is the use of baking soda, or sodium bicarbonate in baking. $NaHCO_3$ is a base. When it reacts with an acid such as lemon juice, buttermilk, or sour cream in a batter, bubbles of carbon dioxide gas are formed from decomposition of the resulting carbonic acid, and the batter “rises.” Baking powder is a combination of sodium bicarbonate, and one or more acid salts that react when the two chemicals come in contact with water in the batter.

Many people like to put lemon juice or vinegar, both of which are acids, on cooked fish (Figure 16.5b). It turns out that fish have volatile amines (bases) in their systems, which are neutralized by the acids to yield involatile ammonium salts. This reduces the odour of the fish, and also adds a “sour” taste that we seem to enjoy.



$$CH_3COOH + NH_2CH_2CH_2CH_2CH_2NH_2 \longrightarrow CH_3COO^- + NH_3^+CH_2CH_2CH_2CH_2NH_2$$

Acetic acid + Putrescine ⟶ Acetate ion + Putrescinium ion

**Figure 16.5b** A neutralization reaction takes place between citric acid in lemons or acetic acid in vinegar, and the bases in the flesh of fish (credit: *Chemistry (OpenStax)*, CC BY 4.0).

Pickling is a method used to preserve vegetables using a naturally produced acidic environment. The vegetable, such as a cucumber, is placed in a sealed jar submerged in a brine solution. The brine solution favours the growth of beneficial bacteria and suppresses the growth of harmful bacteria. The beneficial bacteria feed on starches in the cucumber and produce lactic acid as a waste product in a process called fermentation. The lactic acid eventually increases the acidity of the brine to a level that kills any harmful bacteria, which require a basic environment. Without the harmful bacteria consuming the cucumbers they are able to last much longer than if they were unprotected. A byproduct of the pickling process changes the flavour of the vegetables with the acid making them taste sour.

These cooking practices emerged from Peoples of the past. For example, many of these modern-day preparation and cooking practices steam from traditional Indigenous ingredients and techniques.

## Indigenous Perspective: Indigenous Food Culture

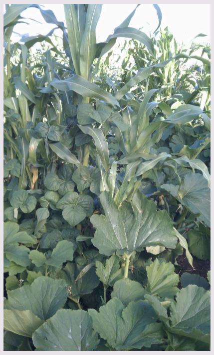


**Figure 16.5c** Three sisters Garden: Three sisters planting (corn, beans, and squash) at a community garden is pictured. (credit: work by esagor, CC BY-NC 2.0.)

Launched in 2020 by the Indigenous Tourism Association of Canada (ITAC), the website, Destination Indigenous, brings awareness and mindfulness to the best Indigenous tourism experiences in Canada from Pacific coast to Arctic coast to Atlantic coast, including Indigenous food culture.

In an excerpt from Food & Culture – Indigenous Cuisine website (2020), food as part of an interdependent ecosystem is highlighted:

> "Traditional food sources were seen as part of a healthy and interdependent ecosystem. Indigenous [communities] traditionally only harvested, hunted or gathered what they needed to survive, and endeavoured to not let anything go to waste. In communities with abundant fish, for example, every edible part of the fish was eaten, including the head, eyes, offal and eggs.
>
> Inedible animal or plant material was often ingeniously repurposed for practical use. Animal bones could be used for tools, tanned hides and furs could be used for shelter and clothing, rawhide could be used for snowshoes, fishing nets or drum covers, and intestines or bladders could be used for cooking vessels or water storage. Plant materials, like spruce root or birch bark, could also be used for food storage." (*Food & Culture – Indigenous Cuisine*, 2020, para. 4).

Neutralization reactions of acids and bases are an important concept in many aspects of life. Conceptualizing the outcomes of acid base reactions is key to understanding the world around us.

## Strong Acid-Strong Base Reactions

When a strong acid and a strong base are combined in stoichiometrically equivalent quantities – when $[H^+]$ equals $[OH^-]$ – a neutral solution results (pH = 7). The acid and base have neutralized each other, and the acidic and basic properties are no longer present. When equal amounts of a strong acid such as hydrochloric acid are mixed with a strong base such as sodium hydroxide, the result is a neutral solution. The products of the reaction do not have the characteristics of either an acid or a base. The balanced molecular equation is:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H2O(l)$$

As discussed in chapter 14, chemical reactions occurring in aqueous solution are more accurately represented with a net ionic equation. The full ionic equation for the neutralization of hydrochloric acid by sodium hydroxide is written as follows:

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$$

Since the acid and base are both strong, they are fully ionized and so are written as ions, as is the NaCl formed as a product. The sodium and chloride ions are spectator ions in the reaction, leaving the following as the net ionic reaction.

$$H^+ (aq) + OH^-(aq) \rightarrow H_2O(l)$$

All neutralization reactions of a strong acid with a strong base simplify to the net ionic reaction of hydrogen ion combining with hydroxide ion to produce water.

**Watch Acid-Base Reaction (HCl + NaOH) (0min 56s) (https://youtu.be/TS-I9KrUjB0).**

What if the acid is a **diprotic acid** such as sulfuric acid? The balanced molecular equation now involves a 1:2 ratio between acid and base.

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$$

In order for the reaction to be a full neutralization, twice as many moles of NaOH must react with the $H_2SO_4$. The sodium sulfate salt is soluble, and so the net ionic reaction is again the same. Different mole ratios occur for other polyprotic acids or bases with multiple hydroxides such as $Ca(OH)_2$.

## Reactions Involving a Weak Acid or Weak Base

Salt solutions do not always neutralize to have a pH of 7. As a general concept, if a strong acid is mixed with a weak base, the resulting solution will be slightly acidic. If a strong base is mixed with a weak acid, the solution will be slightly basic. The reaction between weak nitrous acid and strong potassium hydroxide is shown below.

$$HNO_2(aq) + KOH(aq) \rightarrow KNO_2(aq) + H_2O(l)$$

In order to write the net ionic equation, the weak acid must be written as a molecule since it does not ionize to a great extent in water. The base and the salt are fully dissociated.

$$HNO_2(aq) + K^+(aq) + OH^-(aq) \rightarrow K^+(aq) + NO_2^-(aq) + H_2O(l)$$

The only spectator ion is the potassium ion, resulting in the net ionic equation:

$$HNO_2(aq) + OH^-(aq) \rightarrow NO_2^-(aq) + H_2O(l)$$

The strong hydroxide ion essentially “forces” the weak nitrous acid to become ionized. The hydrogen ion from the acid combines with the hydroxide ion to form water, leaving the nitrite ion as the other product. The resulting solution is not neutral, but instead is slightly basic.

Reactions can also involve a weak base and strong acid, resulting in a solution that is slightly acidic. The molecular and net ionic equations for the reaction of hydrochloric acid and ammonia are shown below.

$$HCl(aq) + NH_3(aq) \rightarrow NH_4Cl(aq)$$

$$H^+(aq) + NH_3(aq) \rightarrow NH_4^+(aq) \text{ (Cl}^-\text{is a spectator ion)}$$

Reactions between acids and bases that are both weak may result in solutions that are neutral, acidic, or basic.

**Source:** "Strong Acid-Strong Base Reactions" was adapted by Jackie MacDonald from "21.16: Neutralization Reaction and Net Ionic Equations for Neutralization Reactions" In *Introductory Chemistry (CK-12)* by CK-12 Foundation, shared under CK-12 license

## Attributions & References

Except where otherwise noted, this page is adapted by Jackie MacDonald from "14.4 Hydrolysis of Salts (https://boisestate.pressbooks.pub/chemistry/chapter/14-4-hydrolysis-of-salts/)" In *General Chemistry 1 & 2* by Rice University, a derivative of *Chemistry (Open Stax)* by Paul Flowers, Klaus Theopold, Richard Langley & William R. Robinson and is licensed under CC BY 4.0. Access for free at *Chemistry (OpenStax) (https://openstax.org/books/chemistry/pages/1-introduction)* / Adaptions and modifications were made to content by for student comprehension.

### References

Food & Culture – Indigenous Cuisine. (2020, September 21). *DI – Culinary*.

# 16.6 TITRATIONS AND NEUTRALIZATION CALCULATIONS

## Learning Objectives

By the end of this section, you will be able to:

- Explain the use of acid-base indicators in determining pH of a solution and outline how it is used in a titration experiment.
- Know the steps of an acid-base titration experiment.
- Perform neutralization calculations involving strong acids and strong bases.

In the previous section, neutralization reactions were introduced. This section will examine the basics of acid-base titrations and perform neutralization calculations between strong acids and strong bases.

## Titrations

When hydrochloric acid is reacted with sodium hydroxide, an acid/base mole ratio of 1:1 is required for full neutralization.

$$\mathrm{HCl}(aq) + \mathrm{NaOH}(aq) \rightarrow \mathrm{NaCl}(aq) + \mathrm{H_2O}(l)$$

If instead the hydrochloric acid were reacted with barium hydroxide, the mole ratio would be 2:1.

$$2\mathrm{HCl}(aq) + \mathrm{Ba(OH)_2}(aq) \rightarrow \mathrm{BaCl_2}(aq) + 2\mathrm{H_2O}(l)$$

Two moles of HCl are required to completely neutralize one mole of $Ba(OH)_2$. In a balanced neutralization equation, the moles of $H^+$ ions supplied by the acid will be equal to the moles of $OH^-$ ions supplied by the base. The equivalence point is the point in a neutralization reaction when the number of moles of hydrogen ions is equal to the number of moles of hydroxide ions.

In the laboratory, you may need to determine the concentration of an acid or a base whose concentration is not known. This can be accomplished by performing a controlled neutralization reaction. A **titration** is an experiment in which a solution, whose concentration is known, is gradually added to a measured volume of another solution in order to determine its concentration. Typically, the **titrant** (the known solution) is added from a buret to a known volume of the **analyte** (the unknown concentration). Many titrations are acid-base neutralization reactions, although other types of titrations can also be performed.

In order to perform an acid-base titration, the chemist must have a way to visually detect that the neutralization reaction has reached the equivalence point. An indicator is a substance that has a distinctly different colour when in an acidic or basic solution. A commonly used indicator for strong acid-strong base titrations is phenolphthalein. Solutions in which a few drops of phenolphthalein have been added, turn from colourless to brilliant pink as the solution turns from acidic to basic (Figure 16.6a).

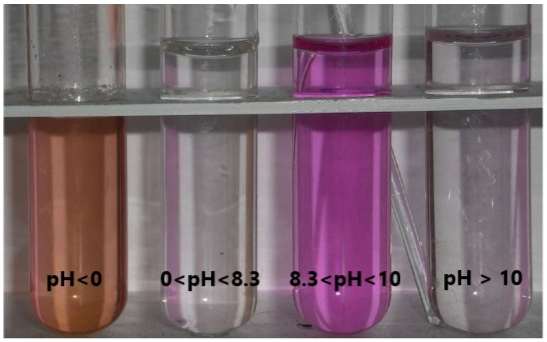


**Figure 16.6a** Phenolphthalein at different pH: Four test tubes are shown containing solutions at various pH, all containing the pH indicator, phenolphthalein. From left to right: At a pH less than zero, phenolphthalein will turn orange, as shown in the first test tube on the left. At a pH between 0 and 8.3, the solution will remain clear. A solution between a pH of zero to 8.3 through 10, the solution will be pink. Finally a solution with this indicator will be clear for any solutions with a pH greater than 10. Exact pH ranges of this indicator are provided by the supplier when purchased used in lab (credit: modification of work by Siegert, PD / pH ranges corrected by Jackie MacDonald).

The steps in a titration reaction are outlined below.

1. A measured volume of an acidic solution (the analyte) whose concentration is unknown is added to an Erlenmeyer flask.
2. Several drops of an indicator are added to the acid and mixed by swirling the flask.
3. A buret is filled with a basic solution (the titrant) of known molarity.
4. The stopcock of the buret is opened and base is slowly added to the acid while the flask is constantly swirled to ensure mixing. The stopcock is closed at the exact point at which the indicator just changes colour. If the colour change does not remain after swirling, the neutralization reaction has not yet reached completion.

The standard solution is the solution in a titration whose concentration is known. In the titration described above, the base solution is the standard solution. It is very important in a titration to add the solution from the buret slowly so that the point at which the indicator changes colour can be found accurately. The end point of a titration is the point at which the indicator changes colour. When phenolphthalein is the indicator, the end point will be signified by a faint pink colour.

Watch **Setting Up and Performing a Titration (6min 52sec)** (https://youtu.be/sFpFCPTDv2w).

Other pH indicators behave in essentially the same way.

**Watch What Are Indicators & How Do We Use Them? | Chemical Tests | Chemistry | FuseSchool (3min 18s (https://youtu.be/xYQlvTblgCY)).**

## Neutralization Calculations

As shown at the beginning of this section, when hydrochloric acid is reacted with sodium hydroxide, an acid/base mole ratio of 1:1 is required for full neutralization as shown in the balance chemical equation below:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

However, as we have learned, not all acids and bases neutralize at a one to one mole ratio. When a strong diprotic or triprotic acid or base is used in a titration neutralization reaction, the mole ratio needs to be taken into consideration when performing calculations. Let's reexamine the neutralization of hydrochloric acid with diprotic barium hydroxide:

$$2HCl(aq) + Ba(OH)_2(aq) \rightarrow BaCl_2(aq) + 2H_2O(l)$$

Two moles of HCl are required to completely neutralize one mole of $Ba(OH)_2$. The strong acid, HCl, furnishes 1 $H^+$ ion; whereas, $Ba(OH)_2$ ionizes 2 $OH^-$ ions. In a balanced neutralization equation, the moles of $H^+$ ions supplied by the acid is equal to the moles of $OH^-$ ions supplied by the base. The equivalence point is the point in a neutralization reaction when the number of moles of hydrogen ions is equal to the number of moles of hydroxide ions.

Therefore, it is of utmost importance that students consider mole ratios of the acid and base in their calculations when performing neutralization calculations.

There are different ways to solve neutralization problems: solving neutralization problems using stoichiometry will be outlined in this section.

### Performing Stoichiometric Calculations to Solve Neutralization Problems

Take a closer look at Figure 16.6b, which is a modified version of Figure 9.2a from 9.2 – Mole-Mass and Mass-Mass Calculations. Focus your attention on the four rectangles inside the red coloured outline. These provide a flowchart on the steps needed to solve a neutralization reaction problem.

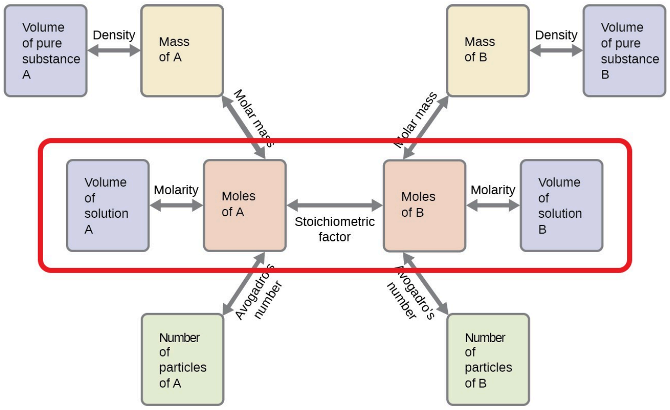


**Figure 16.6b** Stoichiometry Flowchart for Neutralization Calculations. (Image Source: General Chemistry 1 & 2, CC BY 4.0). (credit: *Chemistry (OpenStax)*, CC BY 4.0).

## Calculating the Unknown Molarity of an Acid/Base in a Neutralization Reaction:

Follow the steps below to solve for an unknown molarity.

- Take known volume of known solution "A" and multiply it by molarity of solution A to calculate the moles of A.
- Next, apply the mole ratio calculation to determine the moles of solution B.
- Finally, apply the final stoichiometric calculation using the moles of solution B and given volume of solution B to calculate the molarity of solution B needed to neutralize solution A.

### Example 16.6a

In a titration of hydrochloric acid with barium hydroxide, 32.00 mL of 0.150 *M* HCl is required to neutralize 26.75 mL of the barium hydroxide solution. Calculate the molarity of the barium hydroxide solution. (Remember, you must work in litres for volume since you are performing mole calculations using units of molarity, which are moles/L).

#### Solution

*Step 1: List the known values, write the balanced chemical equation, and and plan the problem.*

Known

- molarity of the HCl solution = 0.150 *M*

- volume of the HCl solution = 32.00 mL = 0.03200L
- volume of the $Ba(OH)_2$ solution = 26.75 mL = 0.02675L
- Balanced Chemical Equation: $2HCl(aq) + Ba(OH)_2(aq) \rightarrow BaCl_2(aq) + 2H_2O(l)$
- Mole ratio of acid to base in balanced chemical equation is 2:1

Unknown

- molarity of the $Ba(OH)_2$ solution = ?

*Step 2 : Determine the moles of HCl that were consumed in the reaction.*

- mol HCl = $M \times L = 0.150 M \times 0.03200 L = 0.00480$ mols of HCl

*Step 3: From the mole ratio, calculate the moles of $Ba(OH)_2$ that reacted.*

- mol $Ba(OH)_2$ = 0.00480 mols HCl x [1 mol $Ba(OH)_2$ / 2 mol HCl] = 0.00240 mols = $2.40 \times 10^{-3}$ mols of $Ba(OH)_2$

*Step 4: Determine the molarity of $Ba(OH)_2$ by dividing the moles by the volume.*

- $[Ba(OH)_2] = 2.40 \times 10^{-3}$ mols / 0.02675 L = 0.0897 *M*

Therefore the molarity of the basic barium hydroxide solution needed to neutralize the acid solution is 0.0897 *M* or $8.97 \times 10^{-2}$ *M*.

**Source:** “Example 16.6a” by Jackie MacDonald, licensed under CC BY-NC 4.0

## Example 16.6b

Suppose that a titration is performed between a strong acid and strong base: 20.70 mL of 0.500 *M* NaOH is required to reach the end point when titrated against 15.00 mL of HBr of unknown concentration. Use the neutralization equation to determine the concentration of hydrobromic acid used in this titration.

### Solution

*Step 1: List the known values, write the balanced chemical equation, and and plan the problem.*

Known

- molarity of the NaOH solution = 0.500 *M*

- volume of the NaOH solution = 20.70 mL
- volume of the HBr solution = 15.00 mL
- Balanced Chemical Equation: HBr(*aq*) + NaOH(*aq*) → NaBr(*aq*) + $H_2O$*(l)*
- Mole ratio of acid to base in balanced chemical equation is 1:1.

Unknown

- molarity of the HBr solution = ?

*Step 2 : Determine the moles of HBr that were consumed in the reaction.*

- mol NaOH = $M \times L = 0.500M \times 20.70L$ = 0.01035 mols of NaOH

*Step 3: From the mole ratio, calculate the moles of HBr that reacted.*

- mol HBr = 0.01035 mols NaOH x [1 mol HBr / 1 mol NaOH] = 0.01035 mols HBr = $1.035 \times 10^{-2}$ mols HBr

*Step 4: Determine the molarity of HBr by dividing the moles by the volume.*

- [HBr] = $1.035 \times 10^{-2}$ mols / 0.015 L = 0.690*M* HBr

Therefore the molarity of the Hydrobromic acid solution needed to neutralize the base solution is 0.690*M*.

**Source:** "Example 16.6b" by Jackie MacDonald, licensed under CC BY-NC 4.0

## Exercise 16.6a

35.00 mL of 0.125 *M* HCl is required to neutralize 25.00 mL of KOH. Determine the concentration of the base solution.

### Check Your Answer[1]

## Exercise 16.6b

15.75 mL of 0.350 *M* $H_2SO_4$ is required to neutralize 10.00 mL of NaOH. Determine the concentration of the sodium hydroxide solution.

### Check Your Answer[2]

**Source:** "Exercise 16.6b" by Jackie MacDonald is licensed under CC BY-NC 4.0.

## Calculating the Unknown Volume of an Acid/Base in a Neutralization Reaction:

If you are asked to solve for an unknown volume of an acid or base in a neutralization reaction problem, follow these steps:

- Take known volume of known solution "A" and multiply it by molarity of solution A to calculate the moles of A.
- Next, apply the mole ratio calculation to determine the moles of solution B.
- Finally, apply the final stoichiometric calculation using the moles of solution B and given molarity of solution B to calculate the volume of solution B needed to neutralize solution A.

## Example 16.6c

What volume, in mL, of a 0.139 *M* solution of barium hydroxide is required to neutralize 45.00 mL of 0.256 *M* HCl?

### Solution

*Step 1: List the known values, write the balanced chemical equation, and plan the problem.*

Known

- molarity of the HCl solution = 0.256 *M*
- volume of the HCl solution = 45.00 mL = 0.04500 L
- molarity of the $Ba(OH)_2$ solution = 0.139 *M*
- Balanced Chemical Equation: $2HCl(aq) + Ba(OH)_2(aq) \rightarrow BaCl_2(aq) + 2H_2O(l)$
- Mole ratio of acid to base in balanced chemical equation is 2:1.

Unknown

- volume of the $Ba(OH)_2$ solution = ?

*Step 2 : Determine the moles of HCl that were consumed in the reaction.*

- mol HCl = *M* × L = 0.256 *M* × 0.04500 L = 0.01152 mols of HCl

*Step 3: From the mole ratio, calculate the moles of $Ba(OH)_2$ that reacted.*

- mol $Ba(OH)_2$ = 0.01152 mols HCl x [1 mol $Ba(OH)_2$ / 2 mol HCl] = 0.00576 mols = $5.76 \times 10^{-3}$ mols of $Ba(OH)_2$

*Step 4: Determine the volume of $Ba(OH)_2$ using moles and molarity of $Ba(OH)_2$.*

- $[Ba(OH)_2]$ = $5.76 \times 10^{-3}$ mols x 1 L / 0.139 mols = 0.0414 L

*Final Step: Convert volume to mL*

- 0.0414 L x 1000 mL/ 1L = 41.4 mL

Therefore 41.4 mL of the barium hydroxide solution is needed to neutralize 45.00 mL of 0.256 *M* HCl.

**Source:** “Example 16.6c” by Jackie MacDonald, licensed under CC BY-NC 4.0

## Exercise 16.6c

What volume, in mL, of a 0.500 *M* solution of sodium hydroxide would be required to neutralize 40.0 mL of a 0.375 *M* $H_2SO_4$ solution?

**Check Your Answer**[3] **Source: "Exercise 16.6c" by Jackie MacDonald, licensed under CC BY-NC 4.0**

## Attribution & References

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## Notes

1. 0.175 *M* KOH
2. 0.276 *M* NaOH
3. *Step 1: List the known values, write the balanced chemical equation, and plan the problem.* Known

    - molarity of the $H_2SO_4$ solution = 0.375 *M*
    - volume of the $H_2SO_4$ solution = 40.0 mL = 0.0400 L
    - molarity of the NaOH solution = 0.500 *M*
    - Balanced Chemical Equation: $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$
    - Mole ratio of acid to base in balanced chemical equation is 1:2.

    Unknown

    - volume of the $H_2SO_4$ solution = ?

    *Step 2 : Determine the moles of $H_2SO_4$ that were consumed in the reaction.*

    - mol $H_2SO_4$ = *M* × L = 0.375 *M* × 0.0400 L = 0.0150 mols of $H_2SO_4$

    *Step 3: From the mole ratio, calculate the moles of NaOH that reacted.*

- mol NaOH = 0.0150 mols $H_2SO_4$ x [2 mol NaOH / 1 mol $H_2SO_4$ ] = 0.0300 mols = 3.00 x $10^{-2}$ mols of NaOH

*Step 4: Determine the volume of NaOH using moles and molarity of NaOH.*

- [NaOH] = 3.00 x $10^{-2}$ mols x 1 L / 0.500 mols = 0.0600 L

*Final Step: Convert volume to mL*

- 0.0600 L x 1000 mL/ 1L = 60.0 mL

Therefore 60.0 mL of the sodium hydroxide solution is needed to neutralize 40.0 mL of a 0.375 *M* $H_2SO_4$ solution.

# 16.7 BUFFERS

## Learning Objectives

By the end of this section, you will be able to:

- Define buffer and describe how it reacts with an acid or a base.

Weak acids are relatively common, even in the foods we eat. But we occasionally come across a strong acid or base, such as stomach acid, that has a strongly acidic pH of 1–2. By definition, strong acids and bases can produce a relatively large amount of hydrogen or hydroxide ions and, as a consequence, have marked chemical activity. In addition, very small amounts of strong acids and bases can change the pH of a solution very quickly. If 1 mL of stomach acid [which we will approximate as 0.05 M HCl(*aq*)] is added to the bloodstream, and if no correcting mechanism is present, the pH of the blood would go from about 7.4 to about 4.9 — a pH that is not conducive to life. Fortunately, the body has a mechanism for minimizing such dramatic pH changes. This mechanism involves a buffer, a solution that resists dramatic changes in pH.

**Watch Buffers, the Acid Rain Slayer: Crash Course Chemistry #31 (11min 40s) (https://youtu.be/8Fdt5WnYn1k).**

Buffers resist dramatic changes in pH by being composed of certain pairs of solutes: either a weak acid plus a salt derived from that weak acid, or a weak base plus a salt of that weak base. For example, a buffer can be composed of dissolved acetic acid ($HC_2H_3O_2$, a weak acid) and sodium acetate ($NaC_2H_3O_2$, a salt derived from that acid). Another example of a buffer is a solution containing ammonia ($NH_3$, a weak base) and ammonium chloride ($NH_4Cl$, a salt derived from that base).

## Characteristics of a Good Buffer

Good buffering systems have the following characteristics:

- The solution contains a weak acid and its conjugate base OR a weak base and its conjugate acid
- The buffer resists changes in pH by reacting with added acid or base, so these ions do not accumulate.
- Any added acid reacts with the conjugate base to resist pH changes
- Any added base reacts with the conjugate acid to resist pH changes

Buffers cannot be made from a strong acid (or strong base) and its conjugate since these solutions ionize completely in water. Also take note, water is not a buffer.

**Source:** "Characteristics of a Good Buffer" by Jackie MacDonald, CC BY-NC-4.0

## How Buffers Work

Let's consider an acetic acid – sodium acetate buffer to demonstrate how buffers work. If a strong base — a source of $OH-(aq)$ ions — is added to the buffer solution, those hydroxide ions will react with the acetic acid in an acid-base reaction:

$$HC_2H_3O_2(aq) + OH^-(aq) \rightarrow H_2O(l) + C_2H_3O_2^-(aq)$$

Rather than changing the pH dramatically by making the solution basic, the added hydroxide ions react to make water, and the pH does not change much.

If a strong acid—a source of $H^+$ ions—is added to the buffer solution, the $H^+$ ions will react with the anion from the salt. Because $HC_2H_3O_2$ is a weak acid, it is not ionized much. This means that if lots of hydrogen ions and acetate ions (from sodium acetate) are present in the same solution, they will come together to make acetic acid:

$$H^+(aq)+C_2H_3O_2^-(aq) \rightarrow HC_2H_3O_2(aq)$$

Rather than changing the pH dramatically and making the solution acidic, the added hydrogen ions react to make molecules of a weak acid.

In chemistry texts and sources, you may have noticed that $H^+$ and $H_3O^+$ are used interchangeably in contexts when the proton donor-acceptor mechanism does not need to be emphasized. Since it is easier to write the $H^+$ proton, chemists often use it to represent acid-base reactions or to explain general concepts in buffering systems. Thus, it is permissible to talk about "hydrogen ions" and use the formula $H^+$ in writing chemical equations as long as you remember that they are not to be taken literally in the context of aqueous solutions.

**Source:** “The Hydronium Ion” by Stephen Lower & Avneet Kahlon In *Acids and Bases in Aqueous Solutions*, licensed under CC BY 3.0.

Figure 16.7a illustrates both actions of the acetic acid – sodium acetate buffer.

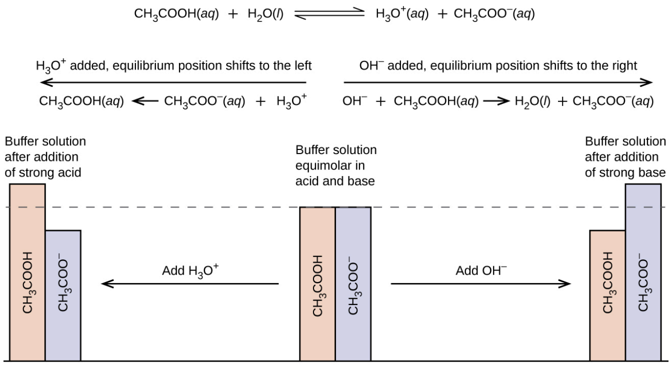

**Figure 16.7a** Buffering Action of Acetic Acid and Sodium Acetate: Whether a small amounts of strong acid or base are added to this buffer system, the pH does not drastically change due to its ability to form more weak acid or more conjugate base depending on whether acid or base is added to the system. (credit: *General Chemistry 1 & 2,* CC BY 4.0),

Figure 16.7b illustrates a basic summary of the action of buffers when small amounts of strong base and acid is added.

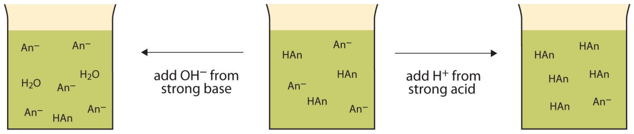

**Figure 16.7b** Generalizing the Actions of Buffers: “HAn” represents the weak Acid. “$An^-$” represents the conjugate base of the weak acid. If strong acid is added to buffer, the system shifts to the right to produce more weak acid, which holds pH steady. If a strong base is added to the buffer, the system shifts to the right forming water and more weak acid, which also holds the pH constant (credit: *Map: Introductory Chemistry (Tro)*, CK-12 license)

Buffers made from weak bases and salts of weak bases act similarly. For example, in a buffer containing $NH_3$ and $NH_4Cl$, ammonia molecules can react with any excess hydrogen ions introduced by strong acids:

$$NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$$

while the ammonium ion, $NH_4^+(aq)$ can react with any hydroxide ions introduced by strong bases:

$$NH_4^+(aq) + OH^-(aq) \rightarrow NH_3(aq) + H_2O(l)$$

## Example 16.7a

Which solute combinations can make a buffer solution? Assume that all are aqueous solutions.

1. $HCHO_2$ and $NaCHO_2$
2. HCl and NaCl
3. $CH_3NH_2$ and $CH_3NH_3Cl$
4. $NH_3$ and NaOH

### Solution

1. Formic acid ($HCHO_2$) is a weak acid, while $NaCHO_2$ is the salt made from the anion of the weak acid—the formate ion ($CHO_2^-$). The combination of these two solutes would make a buffer solution.
2. Hydrochloric acid (HCl) is a strong acid, not a weak acid, so the combination of these two solutes would not make a buffer solution.
3. Methylamine ($CH_3NH_2$) is like ammonia with one of its hydrogen atoms substituted with a $CH_3$ (methyl) group. Because it is not on our list of strong bases, we can assume that it is a weak base. The compound $CH_3NH_3Cl$ is a salt made from that weak base, so the combination of these two solutes would make a buffer solution.
4. Ammonia ($NH_3$) is a weak base, but NaOH is a strong base. The combination of these two solutes would not make a buffer solution.

## Exercise 16.7a

Which solute combinations can make a buffer solution? Assume that all are aqueous solutions.

1. $NaHCO_3$ and NaCl
2. $H_3PO_4$ and $NaH_2PO_4$
3. $NH_3$ and $(NH_4)_3PO_4$
4. NaOH and NaCl

**Check Your Answer**[1]

Buffers work well only for limited amounts of added strong acid or base. Once either solute is all reacted, the solution is no longer a buffer, and rapid changes in pH may occur. We say that a buffer has a certain capacity. Buffers that have more solute dissolved in them to start with have larger capacities, as might be expected.

Human blood has a buffering system to minimize extreme changes in pH. One buffer in blood is based on the presence of $HCO_3^-$ and $H_2CO_3$ [$H_2CO_3$ is another way to write $CO_2(aq)$]. With this buffer present, even if some stomach acid were to find its way directly into the bloodstream, the change in the pH of blood would be minimal. Inside many of the body's cells, there is a buffering system based on phosphate ions.

### Buffered Aspirin

Many people are aware of the concept of buffers from buffered aspirin. Aspirin is well known as a pain reliever and fever reducer. Buffered aspirin contains aspirin (acetylsalicylic acid) and also has magnesium carbonate, calcium carbonate, magnesium oxide, or some other salt. The salt regulates the acidity of the aspirin to minimize its acidic side effects in the stomach. The salt acts like a base, while aspirin is itself a weak acid due to its carboxylic acid group. The H atom in that group can be donated, and therefore, aspirin can act as a Brønsted-Lowry acid. Figure 16.7c and 16.7d show the molecular structure of aspirin in 3D and 2D.

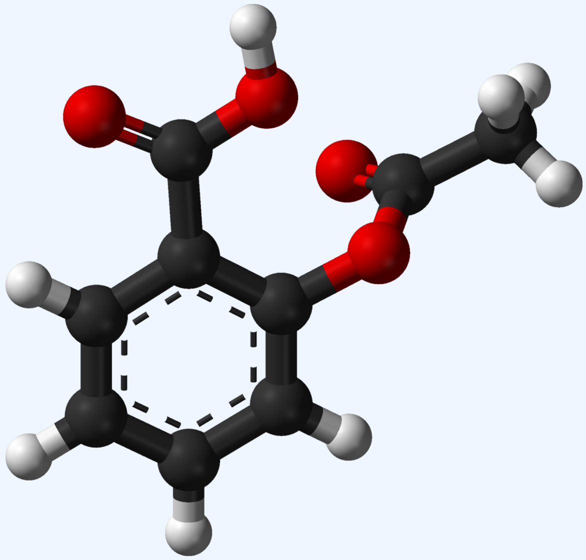

**Figure 16.7c** 3D Ball and Stick Molecular Structure of Aspirin: Aspirin (acetylsalicylic acid) is an aromatic compound containing both a carboxylic acid functional group and an ester functional group rendering it is a weak acid that is only slightly soluble in water. (credit: work by Ben Mills, PD).

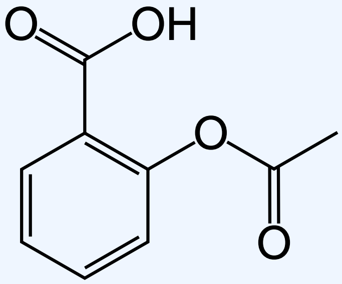

**Figure 16.7d** Skeletal 2D Image of Aspirin: Aspirin (acetylsalicylic acid) is an aromatic compound containing both a carboxylic acid functional group and an ester functional group rendering it is a weak acid. (credit: work by Ben Mills, PD).

## Links to Interactive Learning Tools

Explore Learn the Basics about Buffers (https://h5pstudio.ecampusontario.ca/content/21257) from eCampusOntario H5P Studio (https://h5pstudio.ecampusontario.ca/).

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### References

Stephen Lower & Avneet Kahlon. (2022, August 10). *The Hydronium Ion*. Chemistry LibreTexts.

## Notes

1. 
    1. No; $NaHCO_3$ and NaCl are not acid/base conjugate pairs;
    2. Yes;
    3. Yes; $H_3PO_4$ is a weak acid and $NaH_2PO_4$ is a salt of its conjugate base;
    4. No - NaOH is a strong base, a buffer requires a weak base or acid and its conjugate.

# CHAPTER 16 - SUMMARY

## 16.1 Acids and Bases

A summary of key properties and characteristics of acids and bases commonly encountered in everyday life was provided.

- Both acids and bases release ions in water solution.
- Since both are capable of producing ions when dissolved in water, they are electrolytes.
- Since they are electrolytes, they are both capable of conducting electricity.
- Both acids and bases undergo neutralization reaction

Strong acids and bases are 100% ionized in aqueous solution. Weak acids and bases are less than 100% ionized in aqueous solution. Salts of weak acids or bases can affect the acidity or basicity of their aqueous solutions. A compound that can donate a proton (a hydrogen ion) to another compound is called a Brønsted-Lowry acid. The compound that accepts the proton is called a Brønsted-Lowry base. The species remaining after a Brønsted-Lowry acid has lost a proton is the conjugate base of the acid. The species formed when a Brønsted-Lowry base gains a proton is the conjugate acid of the base. Thus, an acid-base reaction occurs when a proton is transferred from an acid to a base, with formation of the conjugate base of the reactant acid and formation of the conjugate acid of the reactant base. Amphiprotic species can act as both proton donors and proton acceptors. Water is the most important amphiprotic species. It can form both the hydronium ion, H3O+, and the hydroxide ion, OH– depending on what it is reacting with.

## 16.2 Ionization of Water

An acid-base reaction occurs when a proton is transferred from an acid to a base, with formation of the conjugate base of the reactant acid and formation of the conjugate acid of the reactant base. Amphiprotic species can act as both proton donors and proton acceptors. Water is the most important amphiprotic species. It can form both the hydronium ion, $H_3O^+$, and the hydroxide ion, $OH^-$ when it undergoes autoionization:

$$2\mathrm{H_2O}(l) \leftrightharpoons \mathrm{H_3O^+}(aq) \ + \ \mathrm{OH^-}(aq)$$

The ion product of water, $K_\mathrm{w}$ is the equilibrium constant for the autoionization reaction:

$$K_\mathrm{w} = [\mathrm{H_3O^+}][\mathrm{OH^-}] = 1.0 \ \times \ 10^{-14} \text{ at } 25\ ^\circ\mathrm{C}$$

## 16.3 Reactions in Acids and Bases

Chemical reactions are classified according to similar patterns of behaviour. Acid-base reactions involve the transfer of hydrogen ions between reactants.

General acid-base reactions, also called neutralization reactions can be summarized with the following reaction equation:

$$\text{ACID}(aq) + \text{BASE}(aq) \rightarrow H_2O(l) + \text{SALT}(aq) \text{ or } (s)$$

- The DRIVING FORCE for a general acid-base reaction is the formation of water.

Gas-forming acid-base reactions can be with carbonates or metals. One example of an acid mixing with a carbonate is summarized in the following reaction equation:

$$\text{ACID}(aq) + NaHCO_3 \text{ or } Na_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g) + \text{SALT}(aq) \text{ or } (s)$$

- The DRIVING FORCE for a gas-forming acid-base reaction is the formation of gas.

Reactions of acids with metal oxides results in the formation of salt and water.

## 16.4 Introduction to pH and pOH

The concentration of hydronium ion in a solution of an acid in water is greater than 1.0 × $10^{-7}$ $M$ at 25 °C. The concentration of hydroxide ion in a solution of a base in water is greater than 1.0 × $10^{-7}$ $M$ at 25 °C. The concentration of $H_3O^+$ in a solution can be expressed as the pH of the solution; $\text{pH} = -\log[H_3O^+]$. The concentration of $OH^-$ can be expressed as the pOH of the solution: $\text{pOH} = -\log[OH^-]$. In pure water, pH = 7.00 and pOH = 7.00.

## 16.5 Neutralization

The characteristic properties of aqueous solutions of Brønsted-Lowry acids are due to the presence of hydronium ions; those of aqueous solutions of Brønsted-Lowry bases are due to the presence of hydroxide ions. The neutralization that occurs when aqueous solutions of acids and bases are combined results from the reaction of the hydronium and hydroxide ions to form water. A strong acid will neutralize a strong base and result in a neutral solution. However, some salts formed in neutralization reactions may make the product solutions slightly acidic or slightly basic.

## 16.6 Titrations and Neutralization Calculations

Acids and bases neutralize each other, forming a salt and water. A strong acid-strong base neutralization results in a neutral solution with a pH of 7. A titration is an experiment in which a controlled acid-base neutralization reaction is used to determine the unknown concentration of an acid or a base. The equivalence point is reached when the number of hydrogen ions is equal to the number of hydroxide ions. Acid-base indicators are used in a titration in order to detect the end point of the titration. Neutralization calculations can be performed to determine the concentration or volume of a strong acid or base used in a titration.

## 16.7 Buffers

Buffering systems have the following characteristics:

- The solution contains a weak acid and its conjugate base OR a weak base and its conjugate acid
- The buffer resists changes in pH by reacting with added acid or base, so these ions do not accumulate.
- Any added acid reacts with the conjugate base to resist pH changes
- Any added base reacts with the conjugate acid to resist pH changes

Buffers cannot be made from a strong acid (or strong base) and its conjugate since these solutions ionize completely in water. Also recognize that water is not a buffer.

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# CHAPTER 16 - REVIEW

## 16.1 Acids and Bases

1. Write equations that show $NH_3$ as both a conjugate acid and a conjugate base.
   **Check answers:** [1]
2. Write equations that show $H_2PO_4{}^-$ acting both as an acid and as a base.
   **Check answers:** [2]
3. Show by suitable net ionic equations that each of the following species can act as a Brønsted-Lowry acid:
   a. $H_3O^+$
   b. HCl
   c. $NH_3$
   d. $CH_3CO_2H$
   e. $NH_4{}^+$
   f. $HSO_4{}^-$
   **Check answers:** [3]
4. Show by suitable net ionic equations that each of the following species can act as a Brønsted-Lowry acid:
   a. $HNO_3$
   b. $PH_4{}^+$
   c. $H_2S$
   d. $CH_3CH_2COOH$
   e. $H_2PO_4{}^-$
   f. $HS^-$
   **Check answers:** [4]
5. Show by suitable net ionic equations that each of the following species can act as a Brønsted-Lowry base:
   a. $H_2O$
   b. $OH^-$
   c. $NH_3$
   d. $CN^-$
   e. $S^{2-}$
   f. $H_2PO_4{}^-$
   **Check Answers:** [5]
6. Show by suitable net ionic equations that each of the following species can act as a Brønsted-Lowry base:
   a. $HS^-$

- b. $PO_4{}^{3-}$
- c. $NH_2{}^-$
- d. $O^{2-}$
- e. $H_2PO_4{}^-$

**Check Answers:** [6]

7. What is the conjugate acid of each of the following? What is the conjugate base of each?
   - a. $OH^-$
   - b. $H_2O$
   - c. $HCO_3{}^-$
   - d. $NH_3$
   - e. $HSO_4{}^-$
   - f. $H_2O_2$
   - g. $HS^-$
   - h. $H_5N_2{}^+$

   **Check Answers:** For the following the conjugate acid is written first followed by its conjugate base: [7]
8. What is the conjugate acid of each of the following? What is the conjugate base of each?
   - a. $H_2S$
   - b. $H_2PO_4{}^-$
   - c. $PH_3$
   - d. $HS^-$
   - e. $HSO_3{}^-$
   - f. $H_3O_2{}^+$
   - g. $H_4N_2$
   - h. $CH_3OH$

   **Check Answers:** [8]
9. Identify and label the Brønsted-Lowry acid, its conjugate base, the Brønsted-Lowry base, and its conjugate acid in each of the following equations:
   - a. $HNO_3 + H_2O \longrightarrow H_3O^+ + NO_3{}^-$
   - b. $CN^- + H_2O \longrightarrow HCN + OH^-$
   - c. $H_2SO_4 + Cl^- \longrightarrow HCl + HSO_4{}^-$
   - d. $HSO_4{}^- + OH^- \longrightarrow SO_4{}^{2-} + H_2O$
   - e. $O^{2-} + H_2O \longrightarrow 2OH^-$
   - f. $[Cu(H_2O)_3(OH)]^+ + [Al(H_2O)_6]^{3+} \longrightarrow Cu(H_2O)_4]^{2+} + [Al(H_2O)_5(OH)]^{2+}$
   - g. $H_2S + NH_2{}^- \longrightarrow HS^- + NH_3$

   **Check Answer:** [9]

10. Identify and label the Brønsted-Lowry acid, its conjugate base, the Brønsted-Lowry base, and its conjugate acid in each of the following equations:
    a. $NO_2^- + H_2O \longrightarrow HNO_2 + OH^-$
    b. $HBr + H_2O \longrightarrow H_3O^+ + Br^-$
    c. $HS^- + H_2O \longrightarrow H_2S + OH^-$
    d. $H_2PO_4^- + OH^- \longrightarrow HPO_4^{2-} + H_2O$
    e. $H_2PO_4^- + HCl \longrightarrow H_3PO_4 + Cl^-$
    f. $[Fe(H_2O)_5(OH)]^{2+} + [Al(H_2O)_6]^{3+} \longrightarrow [Fe(H_2O)_6]^{3+} + [Al(H_2O)_5(OH)]^{2+}$
    g. $CH_3OH + H^- \longrightarrow CH_3O^- + H_2$

    **Check Answers:** [10]
11. What are amphiprotic species? Illustrate with suitable equations.

    **Check Answers:** [11]
12. State which of the following species are amphiprotic and write chemical equations illustrating the amphiprotic character of these species.
    a. $NH_3$
    b. $HPO_4^-$
    c. $Br^-$
    d. $NH_4^+$
    e. $ASO_4^{3-}$

    **Check Answers:** [12]

## 16.2 Reactions of Acids and Bases

1. The following salts were produced in an acid-base neutralization reaction. Write the formulas and names of the acid and base from which each of these salts are formed.
    a. $NaCl$
    b. $MgCl_2$
    c. $Na_2SO_4$
    d. $NaNO_3$
    e. $K_3PO_4$

    **Check Answers:** [13]
2. When an acid reacts with carbonates what are the characteristic products?
3. Predict the products of a reaction of hydrobromic acid with sodium carbonate: $Na_2CO_3(s) + 2HBr(aq) \rightarrow$

    **Check Answers:** [14]
4. What is the generic formula for the oxidation of metals in acidic solutions?

**Check Answers:** [15]

5. Predict the products of a reaction of hydrochloric acids with magnesium metal.
   **Check Answers:** [16]
6. What is the generic formula for when a metal oxide reacts with an acid?
   **Check Answers:** [17]
7. Predict the products of a reaction of nitric acid with the metal oxide, copper oxide:
   **Check Answers:** [18]

## 16.3 Ionization of Water

1. Write an equation to show the autoionization of water.
   **Check Answers:** [19]
2. Calculate the $[H^+]$ for a solution at 25°C that is $1.0 \text{ x } 10^{-5}$ $M$ $OH^-$. Is this solution acidic, neutral or basic?
   **Check Answers:** [20]
3. Calculate the $[OH^-]$ for a solution at 25°C that is $2.0 \text{ x } 10^{-2}$ $M$ $H^+$. Is this solution acidic, neutral or basic?
   **Check Answers:** [21]
4. Calculate the $[H^+]$ for a solution at 25°C that is $1.0 \text{ x } 10^{-7}$ $M$ $OH^-$. Is this solution acidic, neutral or basic?
   **Check Answers:** [22]
5. Is the self ionization of water endothermic or exothermic? The ionization constant for water ($K_w$) is $2.9 \times 10^{-14}$ at 40°C and $9.3 \times 10^{-14}$ at 60°C.
   **Check Answer:** [23]

## 16.4 Introduction to pH and pOH

1. Explain why a sample of pure water at 40 °C is neutral even though $[H_3O^+] = 1.7 \times 10^{-7} M$. $K_w$ is $2.9 \times 10^{-14}$ at 40 °C.
   **Check Answers:** [24]
2. The ionization constant for water ($K_w$) is $2.9 \times 10^{-14}$ at 40 °C. Calculate $[H_3O^+]$, $[OH^-]$, pH, and pOH for pure water at 40 °C.
   **Check Answers:** [25]
3. The ionization constant for water ($K_w$) is $9.311 \times 10^{-14}$ at 60 °C. Calculate $[H_3O^+]$, $[OH^-]$, pH, and pOH for pure water at 60 °C.
   **Check Answers:** [26]

4. Calculate the pH and the pOH of each of the following solutions at 25 °C for which the substances ionize completely:
    a. 0.200 $M$ HCl
    b. 0.0143 $M$ NaOH
    c. 3.0 $M$ $HNO_3$
    d. 0.0031 $M$ $Ca(OH)_2$

    **Check Answers:** [27]
5. Calculate the pH and the pOH of each of the following solutions at 25 °C for which the substances ionize completely:
    a. 0.000259 $M$ $HClO_4$
    b. 0.21 $M$ NaOH
    c. 0.000071 $M$ $Ba(OH)_2$
    d. 2.5 $M$ KOH

    **Check Answers:** [28]
6. What are the pH and pOH of a solution of 2.0 M HCl, which ionizes completely?
    **Check Answers:** [29]
7. What are the hydronium and hydroxide ion concentrations in a solution whose pH is 6.52?
    **Check Answers:** [30]
8. Calculate the hydrogen ion concentration and the hydroxide ion concentration in a red wine with a pH of 3.500.
    **Check Answers:** [31]
9. Calculate the hydronium ion concentration and the hydroxide ion concentration in lime juice with a pH of 2.00.
    **Check Answers:** [32]
10. The hydronium ion concentration in a sample of rainwater is found to be $1.7 \times 10^{-6}$ $M$ at 25 °C. What is the concentration of hydroxide ions in the rainwater?
    **Check Answers:** [33]
11. The hydroxide ion concentration in household ammonia is $3.2 \times 10^{-3}$ $M$ at 25 °C. What is the concentration of hydronium ions in the solution?
    **Check Answers:** [34]

## 16.5 Neutralization and 16.6 Titrations and Neutralization Calculations

1. 35.00 mL of 0.125 $M$ HCl is required to neutralize 25.00 mL of KOH. Determine the concentration of the base solution.
    **Check Answers:** [35]

2. Suppose that a titration is performed between a strong acid and strong base: 20.70 mL of 0.500 *M* NaOH is required to reach the end point when titrated against 15.00 mL of HCl of unknown concentration. Determine the concentration of hydrochloric acid used in this titration.
   **Check Answers:** [36]
3. What is the concentration of a $Ba(OH)_2$ solution, if 17.25 mL is required to neutralize 19.10 mL of 0.520 *M* HBr?
   **Check Answers:** [37]
4. In a titration of sulfuric acid with sodium hydroxide, 32.20 mL of 0.250 M NaOH is required to neutralize 26.60 mL of the $H_2SO_4$ solution. Calculate the molarity of the sulfuric acid.
   **Check Answers:** [38]
5. What volume of 0.975 *M* NaOH is needed to neutralize 45.0 mL of 0.225 *M* sulfuric acid, $H_2SO_4$?
   **Check Answers:** [39]
6. What volume of 0.202 *M* $HNO_3$ is required to neutralize each of the following solutions?
   a. 15.5 mL of 0.155 *M* NaOH
   b. 25.1 mL of 0.0391 *M* $Ba(OH)_2$
      **Check Answers:** [40]

## 16.7 Buffers

1. Define buffer. What two related chemical components are required to make a buffer?
   **Check Answers:** [41]
2. Can a buffer be made by combining a strong acid with a strong base? Why or why not?
   **Check Answers:** [42]
3. Of the following options (a – d), which combinations of compounds can make a buffer? Assume aqueous solutions.
   a. HCl and NaCl
   b. $HNO_2$ and $NaNO_2$
   c. $NH_4NO_3$ and $HNO_3$
   d. $NH_4NO_3$ and $NH_3$
      **Check Answers:** [43]
4. For each combination in previous question that is a buffer, write the chemical equations for the reactions of the buffer components when a strong acid and a strong base is added.
   **Check Answers:** [44]

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## Notes

1. One example for $NH_3$ as a conjugate acid: $NH_2{}^- + H^+ \longrightarrow NH_3$; as a conjugate base: $NH_4{}^+(aq) + OH^-(aq) \longrightarrow NH_3(aq) + H_2O(l)$
2. One example for $H_2PO_4{}^-$ behaving as an acid (proton donor): $H_2PO_4^-(aq) + OH^-(aq) \rightarrow HPO_4^{2-}(aq) + H_2O(l)$; One example for $H_2PO_4{}^-$ acting as a base (proton acceptor): $H_2PO_4^-(aq) + H_2O(l) \rightarrow H_3PO_4(aq) + OH^-(aq)$
3. (a) $H_3O^+(aq) \longrightarrow H^+(aq) + H_2O(l)$; (b) $HCl(l) \longrightarrow H^+(aq) + Cl^-(aq)$; (c) $NH_3(aq) \longrightarrow H^+(aq) + NH_2{}^-(aq)$; (d) $CH_3CO_2H(aq) \longrightarrow H^+(aq) + CH_3CO_2{}^-(aq)$; (e) $NH_4{}^+(aq) \longrightarrow H^+(aq) + NH_3(aq)$; (f) $HSO_4{}^-(aq) \longrightarrow H^+(aq) + SO_4{}^{2-}(aq)$
4. (a) $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$; (b) $PH_4^+(aq) \rightarrow H^+(aq) + PH_3(aq)$; (c) $H_2S(aq) \rightarrow H^+(aq) + HS^-(aq)$; (d) $CH_3CH_2COOH(aq) \rightarrow H^+(aq) + CH_3CH_2COO^-(aq)$; (e) $H_2PO_4^-(aq) \rightarrow H^+(aq) + HPO_4^{2-}(aq)$; (f) $HS^-(aq) \rightarrow H^+(aq) + S^{2-}(aq)$
5. (a) $H_2O(l) + H^+(aq) \longrightarrow H_3O^+(aq)$; (b) $OH^-(aq) + H^+(aq) \longrightarrow H_2O(l)$; (c) $NH_3(aq) + H^+(aq) \longrightarrow NH_4{}^+(aq)$; (d) $CN^-(aq) + H^+(aq) \longrightarrow HCN(aq)$; (e) $S^{2-}(aq) + H^+(aq) \longrightarrow HS^-(aq)$; (f) $H_2PO_4{}^-(aq) + H^+(aq) \longrightarrow H_3PO_4(aq)$
6. (a) $HS^-(aq) + H^+(aq) \rightarrow H_2S(aq)$; (b) $PO_4{}^{3-} + H^+(aq) \rightarrow HPO_4^{2-}$(aq); (c) $NH_2{}^- + H^+(aq) \rightarrow NH_3(aq)$; (d) $O^{2-}(aq) + H^+(aq) \rightarrow OH^-(aq)$; (e) $H_2PO_4{}^- + H^+(aq) \rightarrow H_3PO_4(aq)$
7. (a) $H_2O$, $O^{2-}$; (b) $H_3O^+$, $OH^-$; (c) $H_2CO_3$, $CO_3{}^{2-}$; (d) $NH_4{}^+, NH_2{}^-$; (e) $H_2SO_4$, $SO_4{}^{2-}$; (f) $H_3O_2{}^+, HO_2{}^-$; (g) $H_2S$; $S^{2-}$; (h) $H_6N_2{}^{2+}$, $H_4N_2$
8. (a) $H_3S^+$, $HS^-$; (b) $H_3PO_4$, $HPO_4^{2-}$; (c) $PH_4^+$, $PH_2^-$; (d) $H_2S$, $S^{2-}$; (e) $H_2SO_3$, $SO_3^{2-}$; (f) $H_4O_2^{2+}$, $H_2O_2$; (g) $H_5N_2^+$, $H_3N_2^-$ ; (h) $CH_3OH_2^+$; $CH_3O^-$ (Watch the video "14.8h | How to find the conjugate acid and conjugate base of

CH3OH (https://youtu.be/9A-DlHNfJo0)" for an explanation of the answer for (h))

9. The labels are Brønsted-Lowry acid = BA; its conjugate base = CB; Brønsted-Lowry base = BB; its conjugate acid = CA. (a) $HNO_3$(BA), $H_2O$(BB), $H_3O^+$(CA), $NO_3{}^-$ (CB); (b) $CN^-$(BB), $H_2O$(BA), HCN(CA), $OH^-$(CB); (c) $H_2SO_4$(BA), $Cl^-$(BB), HCl(CA), $HSO_4{}^-$ (CB); (d) $HSO_4{}^-$ (BA), $OH^-$(BB), $SO_4{}^{2-}$ (CB), $H_2O$(CA); (e) $O^{2-}$(BB), $H_2O$(BA) $OH^-$(CB and CA); (f) $[Cu(H_2O)_3(OH)]^+$(BB), $[Al(H_2O)_6]^{3+}$(BA), $[Cu(H_2O)_4]^{2+}$(CA), $[Al(H_2O)_5(OH)]^{2+}$(CB); (g) $H_2S$(BA), $NH_2{}^-$ (BB), $HS^-$(CB), $NH_3$(CA)

10. The labels are Brønsted-Lowry acid = BA; its conjugate base = CB; Brønsted-Lowry base = BB; its conjugate acid = CA. (a) $NO_2{}^-$(BB), $H_2O$ (BA), $HNO_3$ (CA), $OH^-$(CB); (b) HBr (BA), $H_2O$ (BB), $Br^-$ (CB), $H_3O^+$ (CA); (c) $HS^-$ (BB), $H_2O$ (BA), $H_2S$ (CA), $OH^-$ (CB); (d) $H_2PO_4{}^-$ (BA), $OH^-$ (BB), $HPO_4{}^{2-}$ (CB), $H_2O$ (CA); (e) $H_2PO_4{}^-$ (BB), HCl (BA), $H_3PO_4$ (CA), $Cl^-$ (CB); (f) $[Fe(H_20)_5(OH)]^{2+}$ (BB), $[Al(H_2O)_6]^{3+}$ (BA), $[Fe(H_20)_6]^{3+}$ (CA), $[Al(H_2O)_5(OH)]^{2+}$ (CB); (View the video "14.10f | How to identify the conjugate acid-base pairs in [Fe(H2O)5(OH)]2+ + [Al(H2O)6]3+" for an explanation of the answer for (f): ; (g) $CH_3OH$ (BA), $H^-$ (BB), $CH_3O^-$ (CB), $H_2$ (CA)

11. Amphiprotic species may either gain or lose a proton in a chemical reaction, thus acting as a base or an acid. An example is $H_2O$. As an acid: $H_2O(aq) + NH_3\,(aq) \leftrightharpoons NH_4^+\,(aq) + OH^-\,(aq)$ As a base:
H_{2}O\textit{(aq)} + HCl\textit{(aq)}
\leftrightharpoons H_{3}O^{+}\textit{(aq)} + Cl^{-
}\textit{(aq)}

12. amphiprotic: (a)
$$NH_3 + H_3O^+ \longrightarrow NH_4OH + H_2O$$
$$NH_3 + OCH_3{}^- \longrightarrow NH_2{}^- + CH_3OH$$
(b)
$$HPO_4{}^{2-} + OH^- \longrightarrow PO_4{}^{3-} + H_2O$$
$$HPO_4{}^{2-} + HClO_4 \longrightarrow H_2PO_4{}^- + ClO_4{}^-$$
not amphiprotic: (c) $Br^-$ (d) $NH_4{}^+$ (e) $AsO_4{}^{3-}$

13. (a) HCl, NaOH; (b) HCl, $Mg(OH)_2$; (c) $H_2SO_4$, NaOH; (d) $HNO_3$, NaOH; (e) $H_3PO_4$, KOH

14. $Na_2CO_3(s) + 2HBr(aq) \rightarrow 2NaBr(aq) + CO_2(g) + H_2O(l)$

15. acid + metal → hydrogen + ionic compound

16. $2HCl(aq) + Mg(s) \rightarrow H_2(g) + MgCl_2(aq)$

17. acid + metal oxide → salt + $H_2O(l)$

18. $2HNO_3(aq) + CuO(s) \rightarrow CuNO_3(aq) + 2H_2O(l)$

19. $H_2O(l) + H_2O(l) \leftrightharpoons H_3O^+(aq) + OH^-(aq)$

20. $[H^+] = 1.0 \times 10^{-9}\ M$, solution is basic since $[OH^-] > [H^+]$

21. $[OH^-] = 5.0 \times 10^{-13}\ M$, solution is acidic since $[H^+] > [OH^-]$

22. $[H^+] = 1.0 \times 10^{-7}\ M$, the solution is neutral since $[H^+] = [OH^-]$

23. endothermic, temperature is going up from 40 to 60°C. Water absorbs heat on reactant side to shift to make more product (ionization of water). (View the video "14.14 | Is the self-ionization of water endothermic or exothermic? (https://youtu.be/5Wn_-IKed6M)" for an explanation of this answer

24. In a neutral solution $[H_3O^+] = [OH^-]$. At 40 °C, $[H_3O\ ] = [OH^-] = (2.9 \times 10^{-14})^{1/2} = 1.7 \times 10^{-7}$.

25. In a neutral solution $[H_3O^+] = [OH^-]$. At 40 °C, $[H_3O\ ] = [OH^-] = (2.9 \times 10^{-14})^{1/2} = 1.7 \times 10^{-7}$. pH = pOH = $-\log[1.7 \times 10^{-7}\ M] = 6.77$

26. In a neutral solution $[H_3O^+] = [OH^-]$. At 60 °C, $[H_3O\ ] = [OH^-] = (9.311 \times 10^{-14})^{1/2} = 3.051 \times 10^{-7}$. The pH = pOH = $-\log[3.051 \times 10^{-7} M] = 6.5156$

27. (a) pH = 0.699, pOH = 13.301; (b) pOH = 1.845, pH = 12.155; (c) pH = -0.477; pOH = 15.477; (d) Here 0.0031 $M$ $Ca(OH)_2$ yields 2 $[OH^-]$ ions for every one molecule the base. So, $[OH^-] = 6.2 \times 10^{-3}$ and pOH = 2.21, pH = 11.79

28. (a) pH = 3.587; pOH = 10.413; (b) pOH = 0.68; pH = 13.32; (c) Here 0.000071 $M$ $Ba(OH)_2$ yields 2 $[OH^-]$ ions for every one molecule the base. So, $[OH^-] = 1.42 \times 10^{-4}$ and pOH = 3.85, pH = 10.15 (d) pH = −0.40; pOH = 14.40

29. pH = -0.30, pOH = 14.30
30. $[H_3O^+] = 3.0 \times 10^{-7}$ *M*, $[OH^-] = 3.3 \times 10^{-8}$ *M*
31. $[H_3O^+] = 10^{-3.500} = 3.16 \times 10^{-4}$ *M*, $[OH^-] = 10^{-10.5} = 3.16 \times 10^{-11}$ *M*
32. $[H_3O^+] = 1 \times 10^{-2}$*M*; $[OH^-] = 1 \times 10^{-12}$*M*
33. $[OH^-] = 5.9 \times 10^{-9}$ *M*
34. $[H_3O^+] =3.1 \times 10^{-12}$ *M*
35. $[OH^-] = 0.175$ *M*
36. [HCl] = 0.690 *M*
37. $[Ba(OH)_2] = 0.288$ *M*
38. $[H_2SO_4] = 0.151$ *M*
39. 20.8 mL of NaOH is needed to neutralize the acid.
40. (a) 11.9 mL of $HNO_3$ solution is needed to neutralize the base solution; (b) 9.72 mL of $Ba(OH)_2$ is needed to neutralize the acid solution
41. A buffer is the combination of a weak acid or base and a salt of that weak acid or base and they resist a change in pH upon dilution or upon the addition of small amounts of acid or base.
42. No, buffers cannot be made from a strong acid (or strong base) and its conjugate. This is because they both ionize completely.
43. (a) no; (b) yes; (c) no; (d) yes
44. 3b: when a strong acid is added: $NO_2^- + H^+ \rightarrow HNO_2$; when a strong base is added: $HNO_2 + OH^- \rightarrow NO_2^- + H_2O$; 3d: strong base added: $NH_4^+ + OH^- \rightarrow NH_3 + H_2O$; strong acid: $NH_3 + H^+ \rightarrow NH_4^+$