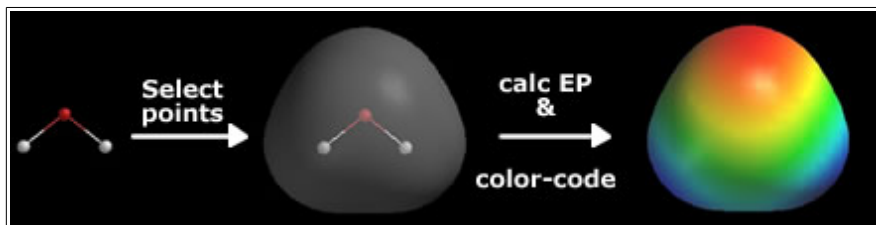


Polar Covalent Bonds

H₂O is a polar covalent molecule. The Bonds between the H atoms and the O atom arise from sharing electrons. These shared electrons form to satisfy the **octet rule**. However, Oxygen is a "selfish" sharer. This **electronegative** aspect of Oxygen means that the electrons of the H₂O molecule preferentially associate near the Oxygen atom, creating partial charges. We indicate this by placing a δ^- near the O and δ^+ 's near the H atoms. These partial charges make the H₂O **polar**.



The electron cloud around a water molecule lingers around the oxygen molecule to render it partially negative. Red illustrates the partial negative end of the molecule while blue indicates the partial positive.

Credit: [FrozenMan \(CC0\)](#)

Because of this polarity, H₂O molecules arrange in a highly structured way. Use the following [simulation to explore polarity of molecules](#).



These weak associations that arise from the polar:polar attractions are referred to as

Hydrogen Bonds (H-bonds). While

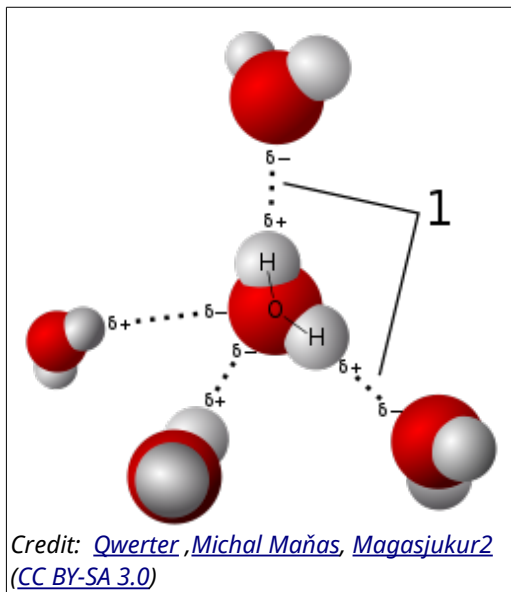
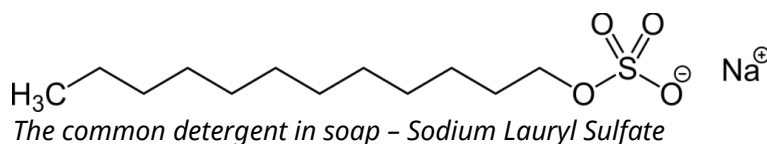
independently weak, the summation of all the H-

bonds are very strong. These associations give rise to the special properties of water: **surface tension**, **cohesion**, **adhesion**, **high specific heat capacity**.

- [Hydrogen Bond Simulation](#)

Polar materials mix with polar materials. Things that can dissolve in water are also polar and referred to as being **hydrophilic** (*hydro-* water, *philic-* liking). Non-polar substances do not interact or mix with polar solvents and are referred to as **hydrophobic** (*hydro-* water, *phobic-* hating). Since carbon and hydrogen share electrons equally, organic compounds are non-polar. Oil is a hydrocarbon that does not mix well with water or vinegar. Vinegar, however is a polar compound that interacts with water. Detergents are called **amphiphilic** (*amphi-* both; *philic-* liking)

because they have portions that are non-polar and portions that are very polar. Detergents can therefore dissolve in hydrocarbons and water. Water alone cannot effectively remove oil from your skin, but a detergent can dissolve the oil and carry it away in water.



Credit: [Qwerter](#), [Michal Mañas](#), [Magasjukur2](#)
(CC BY-SA 3.0)



Surface Tension

Surface tension presents as an invisible film that encompasses the surface of water. The attractive forces arising from the intermolecular **cohesion** holds the surface of water together.



This paperclip would sink if it broke through the surface of the water. Credit: [Alvesgaspar](#) (CC BY-SA 3.0)



These water striders are not on top of the water because they are light. They have not broken through the surface of the water and is therefore, on top of the water. Credit: [Markus Gayda](#) (CC-BY-SA-3.0)

Solutions

Water is an excellent solvent of other polar compounds. Table salt (NaCl) ionizes readily in water. The δ^- O associate around Na^+ while the δ^+ H associate with the Cl^- . If NaCl is dissolved in H_2O , what do you think happens to the intermolecular interactions between water molecules? What do you think would happen to the H-bonds? Would you expect there to be a difference in the surface tension? How do you think this explains the difference of boiling or freezing?

Running on Water



Basiliscus plumifrons (green basilisk)

<https://flic.kr/p/7BTyCX>

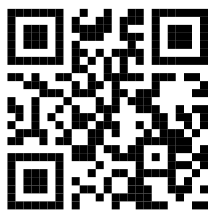


Basiliscus basiliscus (common basilisk)

<https://flic.kr/p/7BXkv7>

Lizards of the genus *Basiliscus* have the nickname "Jesus Christ lizard" for their very special adaptation regarding water.

In the face of danger, these lizards run on their hind legs across water to escape predation. Their hind legs have long toes that help in increasing surface area to distribute their weight so they can propel themselves on the surface of the water. They do not sink because the surface tension of the water is not broken by the large surface covered by their feet. After about 4.5m, they lose sufficient momentum to propel themselves on the water surface and break through. H-bonds enable this adaptation.



<http://youtu.be/45yabrnryXk>

What if?

The namesake of these lizards walked on water but was also able to turn water into wine. Wine is a solution of ethanol (11%) in water. Ethanol has a polar end and a non-polar end.

- How successful would the basilisk run on wine?
- Why?
- How could you test this without necessarily using a lizard?

Acids and Bases

We can call any compound that adds H^+ ions (a free proton) into solution an **acid**. Along with this, we would expect that any compound that would decrease the concentration of free H^+ of a solution as a **base**. **pH** is the power of H^+ of a solution. We define this power as a **molar concentration** of H^+ in solution. This concentration invariably ends up being a relatively small number (though great in absolute numbers) and is expressed as a decimal number. Because the range of the concentrations is so great, we express these numbers as logarithmic numbers to avoid writing many 0's after the decimal and to facilitate communicating the concentration. Since these numbers are so (relatively) small, we use the negative logarithm to describe this concentration.

Mathematically defined,
$$pH = -\log_{10}[H^+]$$

The pH scale ranges so that anything below pH 7 is acidic and anything above pH 7 is alkaline. So a smaller number is more acidic. But didn't we just state that something acidic contains more H^+ ions? Remember, because we are dealing with a negative Logarithm, this means the concentration is higher.

Logarithmic Scales

If we have a quantity that is 10^2 , we know that translates into 100. Just as if we have a quantity of 10^4 , we know that translates into 10000. Just as it becomes inconvenient to keep writing all those 0's, it's really impractical to write many many 0's after a decimal. It's really hard to talk about too! So we likewise will express numbers like 0.0001 as 10^{-4} . A logarithm is the reverse function of an exponent. Therefore:

$$\log_{10}(0.0001) = -4$$

$$\log_{10}(10^{-4}) = -4$$

So how do we define a solution that is pH 2? Well, we already decided that this solution is below pH 7 — making it an acid. But what does this mean in terms of H^+ ion concentration?

Let's work this out algebraically:

$$\bullet pH = -\log_{10}[H^+]$$

Let's bring the (-) over to the other side

$$\bullet -pH = \log_{10}[H^+]$$

Now let's reverse the Log \rightarrow base 10

$$\bullet -pH = \log_{10}[H^+]$$

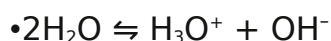
Plug in the pH \rightarrow molar concentration of $[H^+]$

As we can now see, a solution of pH 2 is acidic because the molar concentration of $[H^+]$ is 10^{-2} mole/L or 0.01M

Dissociation of ions: That number is small!

It's not a small number. Remember that a mole is 6.022×10^{23} . That's a very large number! Think about it! A solution of pH 4 is acidic, but if we plug in the formula, we realize that this is equal to 0.0001M H^+ – less than pH 2 at 0.01M!

But let's compare it to the $[H^+]$ content of H_2O . Now I'm going to sound crazier! Water can be thought of as being in an equilibrium where some of the molecules are ionizing and deionizing. We can express this in 2 ways:



So at any given point, a liter of H_2O at neutral pH (7) has 10^{-7} moles of H^+ ions. Incidentally, it also has 10^{-7} moles of OH^- in solution. The second expression indicates the formation of a [hydronium ion \(\$H_3O^+\$ \)](#) instead of a free proton in solution. So something that is pH 2 is a stronger acid than pH 4, right? Nope. That just indicates the amount of free protons in solution. It is more acidic but acid strength means something else. When we talk about **strong acids**, it means that it is more likely to donate a proton to the solution because it is more likely to ionize. Let's look at the following:



Where HA is an acid dissociating in solution

If this dissociation is very high, then we say that it is a strong acid. Similarly, a compound like NaOH readily dissociates completely in solution and provides OH^- ions that can

readily remove H^+ from solution –a strong base! We speak of dissociation in terms of rates and we express this as the [acid dissociation constant](#), K_a . This is calculated using the concentrations of $[H^+]$ (proton), $[A^-]$ (conjugate base) and $[HA]$ (non-dissociated) at equilibrium:

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

Just like the orders of magnitude we have when discussing pH, the rates of dissociation are more conveniently communicated on a logarithmic scale.

$$pK_a = -\log_{10} K_a$$

Think about it this way, if the concentration of the dissociated ions is very high, the numerator in the rate is very high $\rightarrow K_a$ is great. In other words, at equilibrium, the dissociation reaction looks more unidirectional than bi-directional as the compound is readily ionized:

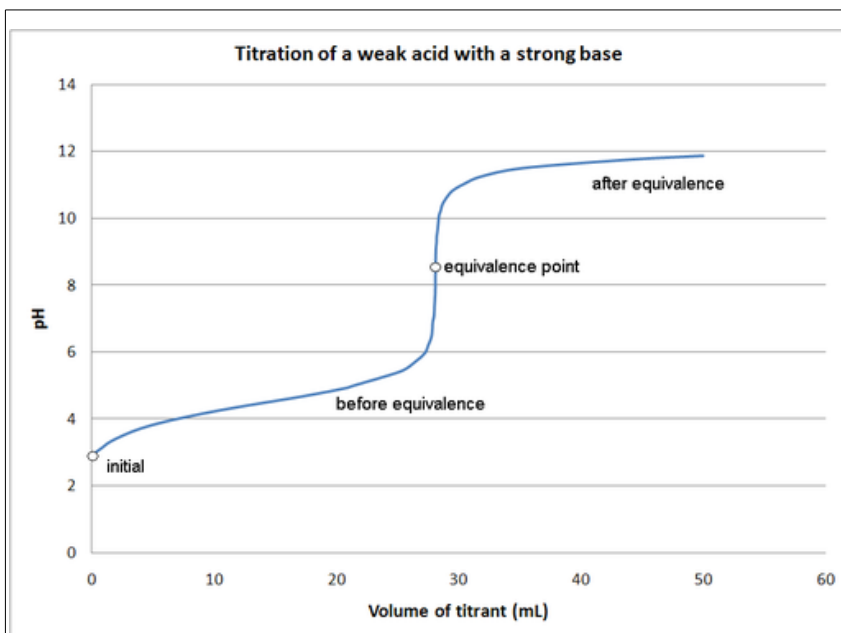


On this scale, we refer to anything with a $pK_a < -2$ as a strong acid since it will readily dissociate in solution. This form of the dissociation constant is extremely useful in estimating the pH of **buffered solutions** and for finding the equilibrium pH of the acid-base reaction (between the proton and the conjugate base). We can estimate the pH by utilizing the [Henderson-Hasselbalch Equation](#):

$$pH = pK_a + \log_{10} \left(\frac{[A^-]}{[HA]} \right)$$

Buffered Solutions

A **buffer** is something that resists change. A **buffered solution** is one that consists of a weak acid or weak base that will control the pH of a solution. Imagine a buffer to be a reservoir of available H^+ or OH^- ions. If a buffered solution is pH 2, adding a basic solution to it will not cause a drastic change to the pH because the reservoir of H^+ will continuously neutralize the base. Eventually, this store or reservoir of H^+ will be depleted. When this happens, the pH will suddenly change. The range in which acid or base is added without a significant change in pH is called the buffered zone or the **buffering capacity**. When this store of H^+ or buffering capacity is expended, we have reached the **equivalence point** that describes the point at which the base has completely neutralized the weak acid.



The solution has a good buffering capacity between pH 3 and pH 5
Credit: [Quantumkinetics](#) (CC BY 3.0)

Acids and Bases Simulation

Acid-Base Solutions

http://phet.colorado.edu/sims/html/acid-base-solutions/latest/acid-base-solutions_en.html

pH simulation

pH Scale

https://phet.colorado.edu/sims/html/ph-scale/latest/ph-scale_en.html

pH of common items

Use the table below to indicate whether an item is acid or base and what you predict the pH to be. We can determine these in class through measuring.

| Solution | Acid or Base | Predicted pH | Actual pH |
|----------------------------|--------------|--------------|-----------|
| Coffee | | | |
| Cola | | | |
| Distilled H ₂ O | | | |
| Detergent | | | |
| Bleach | | | |
| Apple Juice | | | |
| Antacid Solution | | | |

Inquire!

- What is the mechanism of action for antacids?
- Do they all antacids have the same efficacy?

Explore: Determine the pH of common solutions

Use the table below to indicate whether an item is acid or base and what you predict the pH to be. We can determine these in class through measuring. Using the pH meter, measure the pH and validate your predictions

| Solution | Acid or Base | Predicted pH | Actual pH |
|----------------------------|--------------|--------------|-----------|
| Coffee | | | |
| Cola | | | |
| Distilled H ₂ O | | | |
| Detergent | | | |
| Bleach | | | |
| Apple Juice | | | |
| Antacid Solution | | | |

Discovering Buffers

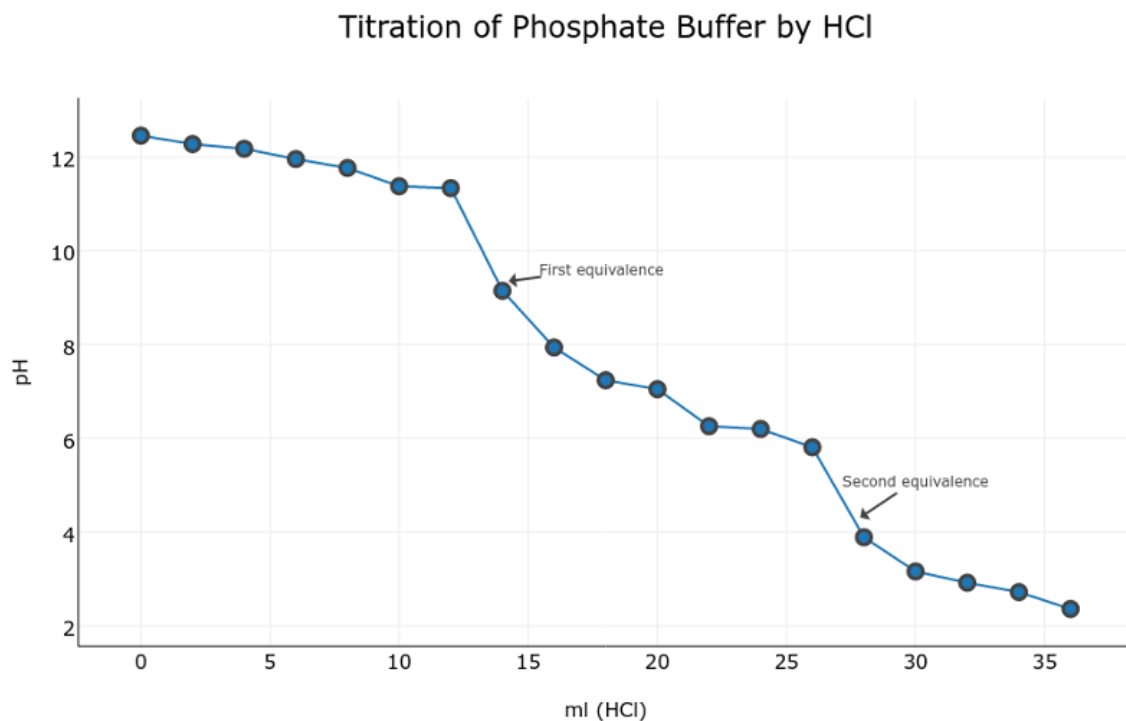
1. Take a small beaker of distilled water and determine its pH. Add one drop of strong acid or of strong base to the water and observe the pH reading on the meter.
2. Now take a small sample of either the standard pH buffer or of any other buffer solution in the lab and add a drop of either strong acid or strong base. Does the reading on the pH meter remain constant or does it change rapidly as it did with the distilled water?

Titration of Acetic Acid

1. Place 50 ml of 0.1M acetic acid into a 250ml beaker
2. Add 5 drops of phenolphthalein indicator and swirl briefly
3. Determine the initial pH with probe
4. Fill a buret with 0.1 N NaOH
5. Add 3 ml volumes NaOH and swirl flask briefly
6. Record the pH and observe the color of the solution
7. Continue to add increments of 3 ml (with swirling) and record the pH until the pH reaches 10.
8. Using a computer, plot pH readings on the y-axis and the volume of base on the x-axis. You can try to use [Plot.ly](https://plot.ly)

Scatterplot Tutorial

| pH | ml (HCl) |
|-------------|-------------|
| choose as x | choose as x |
| choose as y | choose as y |
| 12.46 | 0 |
| 12.28 | 2 |
| 12.18 | 4 |
| 11.96 | 6 |
| 11.77 | 8 |
| 11.38 | 10 |
| 11.34 | 12 |
| 9.15 | 14 |
| 7.94 | 16 |
| 7.24 | 18 |
| 7.05 | 20 |
| 6.26 | 22 |
| 6.2 | 24 |
| 5.81 | 26 |
| 3.89 | 28 |
| 3.16 | 30 |
| 2.92 | 32 |
| 2.72 | 34 |
| 2.36 | 36 |



- [Download the data](#) to try in Plot.ly
- Follow this tutorial on using [Plot.ly](#) to generate a graph. However, do not draw a trendline.
- <https://youtu.be/bxGFS-RukZY>

Titration of Antacid Inquire!

- What pH do you predict an antacid to be?
- Do you expect the antacid to be a strong acid, weak acid, strong base or weak base?
- Does brand of antacid matter in the effectiveness?

Activity

- 1.Retrieve an antacid tablet from the instructor
- 2.Crush the antacid in 100 ml of distilled water
- 3.Transfer 50 ml of the solution to clean 250 ml beaker
 - Share remaining 50 ml with another group (ensure it is a different tablet)
 - if available, add 5 drops of bromocresol purple (0.04% solution) to each sample
- 4.Measure the pH of the antacid solutions and observe color (if using bromocresol purple)
- 5.Determine if you want to titrate with an acid or with a base
- 6.Titrate each antacid solution with either 0.1N HCl or 0.1N NaOH in 2ml increments (you must decide)
 - if using bromocresol purple, stop when yellow is reached

Indicator References

| Bromocresol purple (pH indicator) | |
|-----------------------------------|---------------------|
| <i>below pH 5.2</i> | <i>above pH 6.8</i> |
| 5.2 | 6.8 |

| Phenolphthalein (pH indicator) | |
|--------------------------------|---------------------------------|
| <i>below pH 8.2</i> | <i>between pH 10.0 and 13.0</i> |
| colorless | fuchsia |

Questions for reflection

1. What is a buffer?
2. What is the biological significance of a buffer?
3. Explain differences between strong and weak acids and bases.
4. What are the differences between acids and a bases in terms of $[H^+]$ and $[OH^-]$.
5. Why is pH 7 designated as neutral?
6. How does the antacid work?
7. How much difference in $[H^+]$ is there between pH 4 and pH 8?