EXTENSION ACTIVITY
GENERAL CHEMISTRY

ACID STRENGTH KEY

This activity will serve as practice for the topics covered in the Acid Strength game. This activity is best used in conjunction with not only the tutorial levels, but also supplementary learning resources such as course lectures, textbook reading, etc. Questions labeled “Lock It In” are simply opportunities for you to solidify what you have accomplished in each task and help ensure you meet each objective.

Activity Directions

1. Log into Collisions and navigate to the Acid Strength Game.
2. Play the Tutorial levels, if you haven’t done so already.
3. Exit the levels and enter the Acid Strength sandbox.
4. Follow all instructions as written below. Be sure to reference your course’s textbook, lecture notes, etc. as needed.

OBJECTIVE 1

Demonstrate a basic understanding of acid and base nomenclature.

Since all Brønsted-Lowry acids contain hydrogen as cations, acids are typically named for the anion that they contain.

In **binary acids**—acids that contain only two elements—nomenclature is rather simple. The prefix *hydro-* is added to the root name of the anion and the suffix *-ic* is added to the end along with the word "acid". A classic example is HCl, a binary acid with the name hydrochloric acid. Notice the prefix hydro-, the root name of the chloride ion *-chlor-*, and the suffix *-ic*.

In **oxyacids**—acids that contain hydrogen, oxygen, and another element—the nomenclature also makes use of the anion name. In these acids, anions that end with the suffix *-ate* will have *-ic* at the end of the corresponding acid's name. Anions that end with *-ite* will have *-ous* at the end of the corresponding acid's name. If we consider the chlorate ion (ClO_3^-), its corresponding acid ($HClO_3$) is chloric acid whereas the acid that corresponds to chlorite (ClO_2^-) is chlorous acid. However, the nomenclature gets more complex. If the anion contains more oxygen atoms than the *-ate* ion, it receives the prefix *per-*. This naming convention gives us the *perchlorate ion* (ClO_4^-) and *perchloric acid* ($HClO_4$). If the anion contains one less oxygen than the *-ite* anion, then it receives the prefix *hypo-*. From this naming convention we get the *hypochlorous anion* (ClO^-) and *hypochlorous acid* ($HClO$).

Organic acids are simply organic compounds that have acidic properties. Although they are an important class of acids, some of which appear in this extension activity, the organic naming conventions will not be covered here.

Base nomenclature is less defined than acid nomenclature. That is because many bases are simply ionic compounds that are named using ionic compound nomenclature conventions. The classic examples are the alkali and alkaline earth metal hydroxides. These bases are named like any other ionic compounds (e.g. calcium hydroxide). Other bases typically follow the nomenclature rules of molecular and/or organic compounds.

OBJECTIVE 1

Demonstrate a basic understanding of acid and base nomenclature.

TASK 1: Determine and name the sandbox acids based on the associated anion listed in the table below. The last three are not available in the sandbox, but are good practice.

Anion	Anion Name	Chemical Formula of Acid	Acid Name
F^-	fluoride	HF	Hydrofluoric Acid
Br^-	bromide	HBr	Hydrobromic Acid
I^-	iodide	HI	Hydroiodic Acid
NO_3^-	nitrate	HNO_3	Nitric Acid
SO_4^{2-}	sulfate	H_2SO_4	Sulfuric Acid
CO_3^{2-}	carbonate	H_2CO_3	Carbonic Acid
SO_3^{2-}	sulfite	H_2SO_3	Sulfurous Acid
NO_2^-	nitrite	HNO_2	Nitrous Acid
SO_2^{2-}	hyposulfite	H_2SO_2	Hyposulfurous Acid

LOCK IT IN:
Most of the bases available in the sandbox are ionic compounds and thus follow the relevant nomenclature rules. What is the name of the anion in those compounds? Give the name of one of the ionic bases.

Hydroxide — sodium hydroxide (NaOH)

OBJECTIVE 2

Classify acids as monoprotic or polyprotic based on their chemical and structural formulas.

One way that acids are classified is by the number of protons that are considered 'acidic' or capable of ionizing in solution. Different acids have different numbers of ionizable hydrogens (protons) and thus are deemed **monoprotic** (if only one proton can be produced in solution), **diprotic** (two protons), or triprotic (three protons). Acids with more than one ionizable proton are often just referred to as **polyprotic**.

The method often used to write the chemical formulas of acids not only helps indicate that the substance is an acid, but also the number of protons that can be released by the substance in solution. Since the chemical formula of an acid is often written exactly like an ionic compound, we can simply look at the total number of hydrogen atoms listed first in the name to determine how many ionizable protons an acid will have. The chemical formula of sodium chloride is NaCl where the cation is written first and the anion second such that the overall charge of the formula unit is neutral. We can determine that each formula unit of NaCl will produce one sodium ion and one chloride ion when dissolved in water. Using this same logic, an acid like hydrazoic acid with a chemical formula of HN_3 produces one hydrogen ion (H^+) in solution. It is thus monoprotic.

OBJECTIVE 2

Classify acids as monoprotic or polyprotic based on their chemical and structural formulas.

TASK 2: Determine if the acids in the table are monoprotic, diprotic, triprotic based on their chemical formulas. Almost all of them are available to you in the sandbox.

Chemical Formula	Monoprotic, Diprotic, or Triprotic
H_3PO_4	Triprotic
HCl	Monoprotic
H_2CO_3	Diprotic
HBr	Monoprotic
HCN	Monoprotic
HNO_3	Monoprotic

LOCK IT IN:
The chemical formula for acetic acid can be written in a few different ways, two of which are $C_2H_4O_2$ and $HC_2H_3O_2$. What is the advantage of writing it the second way compared to the first?

Writing acetic acid as $HC_2H_3O_2$ helps indicate that it is a monoprotic acid by placing the acidic hydrogen first. It uses the same formatting as an ionic compound like NaCl where the cation is written before the anion because the hydrogen atom will form a cation in solution. Writing acetic acid as $C_2H_4O_2$ does not readily indicate that it is a monoprotic acid.

OBJECTIVE 2

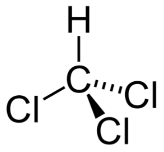
Classify acids as monoprotic or polyprotic based on their chemical and structural formulas.

Determining whether a Brønsted–Lowry acid is monoprotic, diprotic, or triprotic is often possible using the structures of the molecules. The procedures for doing so can be complex and are more suited for higher level chemistry courses. However, there are a few basic patterns that can be used to identify **acidic protons**. In order to do so, it is important to remember that an acidic proton is simply a partial positive hydrogen atom that has a reasonable potential to become a hydrogen ion (H+) in solution. Accordingly, one of the first things to look for in a structure is the presence of hydrogen atoms. However, not all hydrogen atoms are readily ionizable in solution, so we must be careful to identify those which have certain characteristics. A key concept in acid chemistry is that **a proton is more ionizable if the conjugate base that it leaves behind is more stable.** Here are a few guidelines that can help us identify acidic protons based on the stability of the resulting conjugate base:

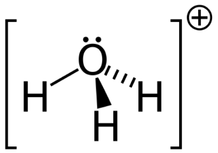
1. Hydrogen atoms bonded to very electronegative atoms are generally more ionizable due to the partial positive character given to the hydrogen. Examples include hydrogen atoms bonded directly to halogens (fluorine, chlorine, bromine, and iodine), oxygen, and nitrogen.

H—Cl

2. Hydrogen atoms attached to carbon atoms and other atoms of similar or lower electronegativity than hydrogen are generally not very acidic. However, highly electronegative atoms attached to carbon (or the other atoms) adjacent to hydrogen will destabilize the bond with hydrogen and stabilize the conjugate base. The result is a more ionizable hydrogen atom. Although chloroform is not typically thought of as an acid, the hydrogen atom in chloroform below is more ionizable than it would be if the electronegative chlorine atoms were not part of the molecule.



3. Hydrogens bonded to electronegative atoms with a positive formal charge *(valence electrons -nonbonding electrons -½ bonding electrons)* are acidic. Take the hydronium ion in the image below. The oxygen atom has a positive formal charge *(6 valence electrons -2 nonbonding electrons -½ x 6 bonding electrons = +1)*. Hydronium will release one proton as a result.

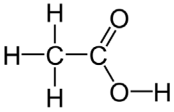
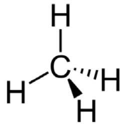
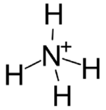
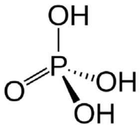
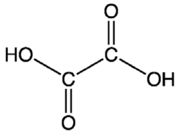


4. There are several other guidelines regarding resonance, orbital hybridization, and other factors. However, these are generally discussed in higher courses and will not be addressed in this activity.

OBJECTIVE 2

Classify acids as monoprotic or polyprotic based on their chemical and structural formulas.

TASK 3: Determine if the substances in the table are acids and then determine if they are monoprotic, diprotic, or polyprotic based on their structures. In the case the substance is not acidic, simply write "N/A" in the third column.

Substance	At Least One Acidic Proton? (Yes or No)	Monoprotic, Diprotic, Triprotic, or N/A
	Yes	Monoprotic
	No	N/A
	Yes	Monoprotic
	Yes	Triprotic
	Yes	Diprotic

LOCK IT IN:
Take a look at the acids available in the sandbox. Find one acid with a hydrogen that is ionizable according to what you learned in the second concept explained for Task 3. Identify the acid and explain why the hydrogen on this acid is less ionizable than the one in HF?

Hydrocyanic acid (HCN) is an example of an acid with a hydrogen that is ionizable according to the second concept in Task 3. The hydrogen in HCN is bonded to the carbon atom, not the nitrogen atom. The electronegativity difference between hydrogen and carbon is not enough that HCN would be acidic for that reason. However, the high electronegativity of the nitrogen atom attached to the carbon helps to destabilize the bond between the carbon and hydrogen enough to make the hydrogen relatively ionizable. The hydrogen in HF, however, is bonded directly to the very electronegative fluorine and is thus more readily ionized.

OBJECTIVE 3

Determine the relative strengths of acids based on their molecular structures.

There are many concepts that underlie the comparison of acid strength, which is defined by how well the acid is able to release hydrogen ions into solution. Much of the basics of these concepts were addressed as part of Task 3. For the purpose of Task 4 and Task 5, we will focus on comparing the strengths of binary acids and oxyacids/oxoacids, both of which can be found in the sandbox.

Binary Acids

Two concepts are central to the strength of binary acids—bond polarity and bond strength. Binary acids have ionizable protons partially because differences in electronegativity between hydrogen and the element to which it is bonded creates a partial positive hydrogen that is more likely to be released as a proton. You will thus notice that binary acids occur when hydrogen is attached to a more electronegative atom. Electronegativity differences alone, however, are not enough to compare the relative strengths of binary acids. **In fact, ranking binary acid strengths by electronegativity differences alone would give you an incorrect ranking.** The relationship between electronegativity and binary acid strength is more nuanced than that.

Bond strength, however, allows us to more accurately rank the strength of binary acids. While the electronegativity difference in a binary acid creates a hydrogen atom that is prepared to ionize, the bond strength will determine just how readily it does so. Release of a hydrogen ion in solution requires the breaking of a bond. A really strong bond requires more energy to break and thus won't happen as readily. In the case of a binary acid, a strong bond will mean that at any one point, only a fraction of ionizable hydrogen atoms will actually be released into the solution. These acids will be less strong. The opposite is true of a weaker bond. Such an acid would more readily release its hydrogen and act as a stronger acid. **Bond strength in binary acids generally decreases with increasing size of the non-hydrogen atom.**

OBJECTIVE 3

Determine the relative strengths of acids based on their molecular structures.

TASK 4: Determine the difference in electronegativity between elements in the binary acids below. Use the electronegativity differences and bond strengths to rank the binary acids in the first table by their strength. Explain your reasoning in the “Justification” column of the table. Check your ranking using the sandbox.

Acid	Electronegativity of Hydrogen	Electronegativity of Second Element	Electronegativity Difference	Bond Strength (kJ/mol)
HI	2.20	2.66	0.46	299
HF	2.20	3.98	1.78	569
H_2S	2.20	2.58	0.38	381
HBr	2.20	2.96	0.76	366
HCl	2.20	3.16	0.96	432

		Justification
Strongest Acid	HI	Hydrogen iodide is both polar and has the lowest bond strength of the substances due to the large size of the iodine atom. The bond very readily breaks in solution and HI is a very strong acid as a result.
	HBr	Hydrogen bromide is both polar and has the second lowest bond strength of the substances listed due to the relatively large size of bromine. It too ionizes very readily in solution.
	HCl	Hydrogen chloride has a greater bond strength than the two stronger acids on the list and thus ionizes less readily. However, it is still a strong acid that ionizes more readily than hydrogen fluoride, which has a greater bond strength.
	HF	Hydrofluoric acid is a weak acid due to the high strength of the bond between hydrogen and the relatively small fluorine. It is the weakest acid of the hydrogen halides as a result.
Weakest Acid	H_2S	Although hydrogen sulfide has a smaller bond strength than hydrogen fluoride, sulfur is much less electronegative than fluorine.

OBJECTIVE 3

Determine the relative strengths of acids based on their molecular structures.

LOCK IT IN:
Use a periodic table to summarize the trend that occurs in the strength of the binary acids formed as you go down the halogen group.

As one goes down the halogen group, the binary acids increase in strength.

LOCK IT IN:
The hydrogen atoms in HF are more acidic than those in H_2O, which are more acidic than those in NH_3. Use a periodic table to identify the trend that occurs in acid strength for hydrides moving from left to right across a period. What explains this trend?

Moving from left to right across a period, the acidity of hydrides increases due to the increase in electronegativity of the non-hydrogen element.

LOCK IT IN:
Check your rankings in the sandbox if you have not already done so. It is quite likely that you put hydrogen sulfide (H_2S) in the incorrect position. Explain how electronegativity can be used to explain the actual positioning of H_2S in the rankings.

Because hydrogen sulfide has a smaller bond strength than hydrogen fluoride it could be reasonably assumed that it is a stronger acid. However, fluorine is much more electronegative than sulfur and thus better stabilizes the resulting negative charge.

OBJECTIVE 3

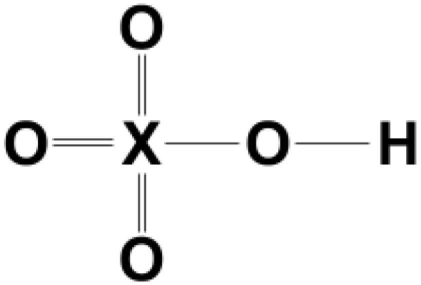
Determine the relative strengths of acids based on their molecular structures.

Oxyacids/Oxoacids

Remember that oxyacids are those that contain a hydrogen atom bonded to an oxygen atom that also contains a third element. A basic structure for an oxyacid is shown in the image below where X represents a third element.

Comparing the relative strengths of oxyacids relies on considering the electronegativity of the third element (X) and the number of oxygen atoms that are bonded to X. As the electronegativity of X increases, the more it helps to pull electrons away from the hydrogen atom and destabilize the oxygen-hydrogen bond. This means that as the electronegativity of X increases, the strength of the acid increases.

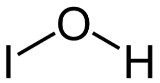
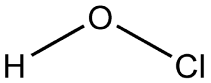
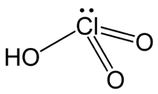
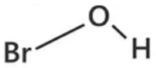
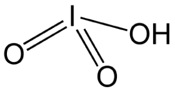
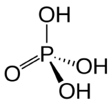
As the number of oxygens bonded to X increases, so does the strength of the acid. This is because oxygen is a very electronegative element that is capable of enhancing the pull of electrons away from the hydrogen atom so that it is released more easily. The acid in the image below has three additional oxygen atoms attached to X compared to our first oxyacid. As a result, the acid below is significantly stronger. Each additional oxygen atom strengthens the acid in a cumulative manner. In the case that X is a very electronegative atom, the acid would be very strong. In cases where X is not very electronegative, even the additional oxygen atoms might not be enough to make the acid strong.



OBJECTIVE 3

Determine the relative strengths of acids based on their molecular structures.

TASK 5: Rank the oxyacids below by their strength. Explain your reasoning in the “Justification” column of the table.

Acid Name and Structure	Chemical Formula	Electronegativity of Non-Hydrogen Element Attached to Oxygen
 Hypoiodous Acid	HIO	Iodine — 2.66
 Hypochlorous Acid	$HClO$	Chlorine — 3.16
 Chloric Acid	$HClO_3$	Chlorine — 3.16
 Hypobromous Acid	HBrO	Bromine — 2.96
 Iodic Acid	HIO_3	Iodine — 2.66
 Phosphoric Acid	H_3PO_4	Phosphorus — 2.19

OBJECTIVE 3

Determine the relative strengths of acids based on their molecular structures.

		Justification
Strongest Acid	Chloric Acid ($HClO_3$)	Not only is the chlorine atom of chloric acid highly electronegative, but the chlorine atom is also bonded to three oxygen atoms. This combination makes chloric acid the strongest of those listed in this activity.
	Iodic Acid (HIO_3)	Although iodine is less electronegative than chlorine, it is still relatively electronegative. In the case of iodic acid, the iodine is bonded to three oxygen atoms, which increases its strength.
	Phosphoric Acid (H_3PO_4)	Phosphoric acid has four oxygen atoms attached to the central phosphorus atom. Although phosphoric acid has more oxygen atoms than iodic or chloric acids, phosphorus has a much lower electronegativity than iodine or chlorine. As a result it is weaker than both of those acids.
	Hypochlorous Acid ($HClO$)	Hypochlorous acid is weaker than the previous three acids due to having just one attached oxygen atom. However, it is stronger than hypobromous or hypoiodous acids due to the higher electronegativity of chlorine.
	Hypobromous Acid ($HBrO$)	Hypobromous acid has only one attached oxygen atom with the electronegativity of bromine being lower than chlorine and higher than iodine.
Weakest Acid	Hypoiodous Acid (HIO)	Hypoiodous acid has only one attached oxygen atom with the electronegativity of iodine being lower than that of both chlorine and bromine. It is the weakest of the acids on the list.

LOCK IT IN:
There are three oxyacids in the sandbox. What are they? Which two would you identify as strong and which one as weak? Explain why that is.

The three oxyacids in the sandbox are sulfuric acid (H_2SO_4), nitric acid (HNO_3), and carbonic acid (H_2CO_3). Sulfuric acid and nitric acid are considered strong acids and carbonic acid is considered a weak one. The main factor that makes sulfuric acid and nitric acid stronger acids is the fact that nitrogen and sulfur are more electronegative than carbon. Sulfuric acid also has an additional oxygen atom, which increases the strength of the acid.

OBJECTIVE 4

Demonstrate an understanding of the acid ionization constant (K_a) and pH.

At this point we have seen some of the ways in which the hydrogen atoms become acidic and how the strengths of acids can be compared structurally. However, one of the most important concepts in acid-base chemistry is a quantitative expression of the strength of an acid known as the **acid ionization constant (K_a).** As you might remember from the LeChâtelier Game, *K* is the equilibrium constant—a ratio used to quantify the relationship between the concentrations or partial pressures of reactants and products at equilibrium. The acid ionization constant is simply the equilibrium constant for the ionization reaction of an acid. As such, the same principles of the equilibrium constant (e.g. an equilibrium constant is for a certain temperature) still apply and the expression is formatted the same way. Remember that the hydrogen ion given off by an acid in solution is transferred to a water molecule and creates hydronium (H_3O^+). As a result, $[H^+] = [H_3O^+]$.

Acid Ionization: $\mathrm{HA\ (aq) + H_2O\ (\ell) \rightleftharpoons H_3O^+ + A^-\ (aq)}$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

Just like you have likely already learned about equilibrium constants for other reactions, the value of K_a can be quite small or quite large. A very large K_a tells us that the forward reaction is favored and thus more hydrogen ions will be produced in solution. A small K_a tells us that the reverse reaction (the one forming the acid) is more favored. Such an acid would not produce as many hydrogen ions in solution. What is often used to discuss the extent to which an acid ionizes is the pK_a ($-\log_{10} K_a$). Very small and increasingly negative pK_a values indicate strong acids. Very large pK_a values indicate very weak acids from which hydrogen ions infrequently appear in solution. Although there is no formal cutoff for weak and strong acids, strong acids typically have pK_a values of less than 1.

OBJECTIVE 4

Demonstrate an understanding of the acid ionization constant (K_a) and pH.

TASK 6: For each sandbox acid listed in the table, write out the expression necessary to calculate the acid ionization constant. For any polyprotic acid, simply write the expressions for **EACH** acidic hydrogen.

Acid and pK_a	Acid Ionization Constant Expression
H_2SO_4 1st $pK_a = -2.8$ 2nd $pK_a = 1.99$	$K_{a1} = \frac{[H_3O^+][HSO_4^-]}{[H_2SO_4]}$ $K_{a2} = \frac{[H_3O^+][SO_4^{2-}]}{[HSO_4^-]}$
HCl $pK_a = -5.9$	$K_a = \frac{[H_3O^+]\,[Cl^-]}{[HCl]}$
HNO_3 $pK_a = -1.4$	$K_a = \frac{[H_3O^+][NO_3^-]}{[HNO_3]}$
HCN $pK_a = 9.21$	$K_a = \frac{[H_3O^+]\,[CN^-]}{[HCN]}$
H_2S 1st $pK_a = 7.0$ 2nd $pK_a \approx 17$	$K_{a1} = \frac{[H_3O^+]\,[HS^-]}{[H_2S]}$ $K_{a2} = \frac{[H_3O^+]\,[S^{2-}]}{[HS^-]}$

LOCK IT IN:
What does the difference between the first and second pKa of the diprotic acids tell us about the relative acidity of the first and second ionizable hydrogens?

The pK_a of the first ionizable hydrogen of both H_2SO_4 and H_2S is much lower than that of the second ionizable hydrogen. This tells us that the first hydrogen is more readily ionized than the second.

Demonstrate an understanding of the acid ionization constant (K_a) and pH.

LOCK IT IN:
Use the pK_a to rank each acid in the table by its strength.

Strongest Acid	HCl
	H_2SO_4
	HNO_3
	H_2S
Weakest Acid	HCN

As you should know from your textbook reading or lecture notes, **pH** is a method of determining the acidity of a solution by taking the $-\log_{10} [H_3O^+]$ of a solution. In the case of a strong acid, $[H_3O^+]$ is equivalent to the concentration of the acid itself since in most circumstances, 100% of the acid molecules ionize in solution. However, since weak acids only partially ionize, a more involved process is required for determining the pH of a solution containing them. The following example shows you how to use ICE tables to determine the pH and percent ionization of a monoprotic weak acid solution available in the sandbox. Reference your textbook, notes, or the LeChâtelier Game extension activity if you are not familiar with ICE tables. Also, please note that the contribution of H_3O^+ by the autoionization of water is negligible except in extremely dilute acid solutions. As such, this process has been ignored in the example.

OBJECTIVE 4

Demonstrate an understanding of the acid ionization constant (K_a) and pH.

Example 1: Find the pH and percent ionization of a 0.500 M HF solution.

Use the balanced ionization equation of the acid to set up an ICE table. Notice that the initial [H_3O^+] has been simplified to 0.00.

$HF\ (aq) + H_2O\ (\ell) \rightleftharpoons H_3O^+\ (aq) + F^-\ (aq)$

	[HF]	[H_3O^+]	[F^-]
Initial	0.500	0.00	0.00
Change			
Equilibrium			

Use x to represent the stoichiometric changes in the concentrations of all reactants and products. Then sum all columns to complete the equilibrium concentration row.

$HF\ (aq) + H_2O\ (\ell) \rightleftharpoons H_3O^+\ (aq) + F^-\ (aq)$

	[HF]	[H_3O^+]	[F^-]
Initial	0.500	0.00	0.00
Change	−x	+x	+x
Equilibrium	0.500−x	0.00 +x	0.00 +x

OBJECTIVE 4

Demonstrate an understanding of the acid ionization constant (K_a) and pH.

<table>
<tr>
<td>Substitute the values from the equilibrium row of the table into the acid ionization constant expression. K_a for HF at 25°C is 6.6×10^{-4}.

Solve the quadratic equation for x. Only the positive root can be valid!</td>
<td>$$K_a = \frac{[H_3O^+]\,[F^-]}{[HF]}$$ $$6.6 \times 10^{-4} = \frac{x^2}{0.500 - x}$$ $$x = \cancel{-0.018} \qquad x = 0.018$$</td>
</tr>
<tr>
<td>Use the determined value of x to determine the equilibrium concentration of $[H_3O^+]$.</td>
<td>
<table>
<tr><th></th><th>[HF]</th><th>$[H_3O^+]$</th><th>$[F^-]$</th></tr>
<tr><td>Initial</td><td>0.500</td><td>0.00</td><td>0.00</td></tr>
<tr><td>Change</td><td>−0.018</td><td>+0.018</td><td>+0.018</td></tr>
<tr><td>Equilibrium</td><td>0.500−0.018 = 0.482</td><td>0.00 +0.018 = 0.018</td><td>0.00 +0.018 = 0.018</td></tr>
</table>
</td>
</tr>
<tr>
<td>Use

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

to determine the pH of the solution.</td>
<td>$$pH = -\log_{10} (0.018) = 1.74$$</td>
</tr>
<tr>
<td>Calculate the percent ionization using the following formula:

$$\% \text{ ionization} = \frac{[H_3O^+]_{equil}}{[HA]_{equil}} \times 100\%$$</td>
<td>$$\% \text{ ionization} = \frac{0.018 \text{ M}}{0.482 \text{ M}} \times 100\% = 3.73\%$$</td>
</tr>
</table>

TASK 7: Use your understanding of acid dissociation to determine the missing values in the table for the acids from the sandbox. Use the sandbox to determine if each is a strong or weak acid, and use the space beneath the table for any math and/or ICE tables.

Acid	Strong or Weak?	K_a	pK_a	Acid Concentration (M)	$[H_3O^+]$	pH	% Ionization
HI	Strong	3.2×10^{9}	–9.5	0.500	0.500	0.301	100%
HCN	Weak	6.2×10^{-10}	9.2	0.05	3.1×10^{-11}	5.3	0.0011%
HCl	Strong	1.3×10^{6}	–6.1	0.10	0.10	1.0	100%
HF	Weak	6.6×10^{-4}	3.2	1.2	0.028	1.6	2.4%

HI

$[HI] = [H_3O^+] = 0.500\ M$

$pH = -\log_{10}[0.500] = 0.301$

HCN

$$HCN\ (aq) + H_2O\ (\ell) \rightleftharpoons H_3O^+\ (aq) + CN^-\ (aq)$$

	$[HCN]$	$[H_3O^+]$	$[CN^-]$
Initial	0.05	0.00	0.00
Change	$-x$	$+x$	$+x$
Equilibrium	$0.05-x$	$0.00+x$	$0.00+x$

OBJECTIVE 4

Demonstrate an understanding of the acid ionization constant (K_a) and pH.

$$K_a = \frac{[H_3O^+][CN^-]}{[HCN]}$$

$$6.2 \times 10^{-10} = \frac{x^2}{0.05 - x}$$

$$x = -\cancel{5.6 \times 10^{-6}} \qquad x = 5.6 \times 10^{-6}$$

$$pH = -\log_{10}(5.6 \times 10^{-6}) = 5.3$$

$$\%\ \text{ionization} = \frac{[H_3O^+]_{equil}}{[HA]_{equil}} \times 100\%$$

$$\%\ \text{ionization} = \frac{5.6 \times 10^{-6}\ M}{0.4999944\ M} \times 100\% = 0.0011\%$$

HCl

$[HCl] = [H_3O^+] = 0.10\ M$
$pH = -\log_{10}[0.10] = 1.0$

OBJECTIVE 4

Demonstrate an understanding of the acid ionization constant (K_a) and pH.

HF

$$HF\ (aq) + H_2O\ (\ell) \rightleftharpoons H_3O^+\ (aq) + F^-\ (aq)$$

	[HF]	$[H_3O^+]$	$[F^-]$
Initial	1.2	0.00	0.00
Change	−x	+x	+x
Equilibrium	1.2−x	0.00 +x	0.00 +x

$$K_a = \frac{[H_3O^+]\,[F^-]}{[HF]}$$

$$6.6 \times 10^{-4} = \frac{x^2}{1.2 - x}$$

$$x = \cancel{-0.028} \qquad x = 0.028$$

$$pH = -\log_{10}(0.028) = 1.6$$

$$\%\ \text{ionization} = \frac{[H_3O^+]_{equil}}{[HA]_{equil}} \times 100\%$$

$$\%\ \text{ionization} = \frac{0.028\ M}{1.172\ M} \times 100\% = 2.4\%$$

OBJECTIVE 5

Demonstrate an understanding of the basic ionization constant (K_b) and pH.

Just like acids, there is an equilibrium constant associated with bases known as the **base ionization constant** (K_b). The basis of K_b is the equation for the ionization of water by a weak base to produce hydroxide ions (OH^-) in solution.

Base Ionization of Water: $B\ (aq) + H_2O\ (\ell) \rightleftharpoons BH^+\ (aq) + OH^-\ (aq)$

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

A larger value of K_b indicates that the forward reaction is favored and thus leads to greater production of hydroxide ions. A large K_b indicates a stronger base. A smaller value of K_b indicates that the reverse reaction is favored and thus fewer hydroxide ions are produced. Comparable to acids, the pK_b is often used to discuss the strength of a base. It is determined using the formula $pK_b = (- \log_{10} K_b)$. The lower pK_b, the stronger the base.

As you should have seen in the Acid Strength Game, the ion created when you remove a proton from an acid is known as the **conjugate base**. The acid HA has the conjugate base A^-. An acid with a very large pK_a is strong partially because it creates a very stable conjugate base. The stability of this conjugate base means that it infrequently removes protons from solution and is thus a weak base. **All strong acids have weak conjugate bases.** As the strength of the acid decreases, the strength of its conjugate base generally increases. However, the conjugate bases of a weak acid is not a strong base. The conjugate base of a weak acid is also a weak base, but stronger than the conjugate bases of strong acids.

The relationship between K_a and K_b is very important. However, the relationship is united through the autoionization constant of water K_w.

Autoionization of Water: $2\ H_2O\ (\ell) \rightleftharpoons H_3O^+\ (aq) + OH^-\ (aq)$

$$K_w = [H_3O^+][OH^-]$$

For a conjugate acid-conjugate base pair, the following is true at just under 25°C:

$$\mathbf{K_w = K_a \times K_b = 1.00 \times 10^{-14}}$$

$$\mathbf{pK_a + pK_b = 14}$$

OBJECTIVE 5

Demonstrate an understanding of the basic ionization constant (K_b) and pH.

TASK 8: For each sandbox base listed below, write out the expression necessary to calculate the base ionization constant. Then use the pKa of the base to determine its pKb. The first two bases are available in the sandbox.

Base and pK_a	Base Ionization Constant Expression	pK_b
Ammonia NH_3 pK_a of NH_4^+ = 9.25	$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$	pK_b = 4.75
Pyridine C_5H_5N pK_a of $C_5H_6N^+$ = 5	$K_b = \frac{[C_5H_6N^+][OH^-]}{[C_5H_5N]}$	pK_b = 9
Carbonate CO_3^{2-} pK_a of HCO_3^- = 10.3	$K_b = \frac{[HCO_3^-][OH^-]}{[CO_3^{2-}]}$	pK_b = 3.7

LOCK IT IN:
Rank the three bases in the table by their strength.

Strongest Base	CO_3^{2-}
	NH_3
Weakest Base	C_5H_5N

LOCK IT IN:
Which of the acids in the table from **Task 7** likely has the strongest conjugate base? The weakest? Explain how you know.

The acid in Task 7 with the strongest conjugate base is HCN. Generally, the weaker the acid, the stronger the conjugate base and vice versa. HCN is the weakest acid in Task 7 as shown by it having the largest pKa of 9.2. This also tells us that the conjugate base (CN^-) has the smallest pK_b (14 - 9.2 = pK_b of 4.8), which reinforces the idea that it is indeed the strongest base. The acid in Task 7 with the weakest conjugate base would be HI since it is the strongest acid (as shown by having the lowest pK_a).

A strong base is a metal hydroxide and dissociates completely in an aqueous solution. As a result, the [OH^-] is the same as the concentration of a strong base. Knowing the concentration of [OH^-] then allows one to determine [H_3O^+] using K_w.

Determining the pH of a weak base solution, however, is very similar to finding the pH of a weak acid solution. The following example will show you how to find the pH of a solution made with a base available in the sandbox. The same assumption about the autoionization of water made in the weak acid example is also used here.

Example 1: Find the pH and percent ionization of a 0.500 M NH_3 solution.

Use the balanced ionization equation of the acid to set up an ICE table. Notice that the initial [OH^-] has been simplified to 0.00.

$$NH_3\,(aq) + H_2O\,(\ell) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)$$

	[NH_3]	[NH_4^+]	[OH^-]
Initial	0.500	0.00	0.00
Change			
Equilibrium			

Demonstrate an understanding of how to interpret heating curves and phase diagrams.

Use x to represent the stoichiometric changes in the concentrations of all reactants and products. Then sum all columns to complete the equilibrium concentration row.

$NH_3\ (aq) + H_2O\ (\ell) \rightleftharpoons NH_4^+\ (aq) + OH^-\ (aq)$

	$[NH_3]$	$[NH_4^+]$	$[OH^-]$
Initial	0.500	0.00	0.00
Change	−x	+x	+x
Equilibrium	0.500−x	0.00 +x	0.00 +x

Plug the values from the equilibrium row of the table into the base ionization constant expression. K_b for NH_3 at 25°C is 1.77×10^{-5}.

Solve the quadratic equation for x. Only the positive root can be valid!

$$K_b = \frac{[BH]\,[OH^-]}{[B]}$$

$$1.77 \times 10^{-5} = \frac{x^2}{0.500 - x}$$

$$x = -\cancel{0.00298} \qquad x = 0.00297$$

<table>
<tr>
<td>Use the determined value of x to determine $[OH^-]$ at equilibrium.</td>
<td>
<table>
<tr><th></th><th>$[NH_3]$</th><th>$[NH_4^+]$</th><th>$[OH^-]$</th></tr>
<tr><td>Initial</td><td>0.500</td><td>0.00</td><td>0.00</td></tr>
<tr><td>Change</td><td>−0.00297</td><td>+0.00297</td><td>+0.00297</td></tr>
<tr><td>Equilibrium</td><td>0.500−0.00297 = 0.49703</td><td>0.00 +0.00297 = 0.00297</td><td>0.00 +0.00297 = 0.00297</td></tr>
</table>
</td>
</tr>
<tr>
<td>Substitute the calculated $[OH^-]$ into the expression for K_w to solve for $[H_3O^+]$
$K_w = [H_3O^+] [OH^-]$
to determine the pH of the solution.</td>
<td>$1.00 \times 10^{-14} = [H_3O^+] (0.00297)$
$[H_3O^+] = 3.37 \times 10^{-12}$</td>
</tr>
<tr>
<td>Use
$pH = -\log_{10} [H_3O^+]$
to determine the pH of the solution.</td>
<td>$pH = -\log_{10} (3.37 \times 10^{-12}) = 11.47$</td>
</tr>
</table>

OBJECTIVE 5

Demonstrate an understanding of how to interpret heating curves and phase diagrams.

TASK 9: Use your understanding of pH to determine the missing values in the table of weak bases. Use the space beneath the table for any math and/or ICE tables.

Base	**K_b**	**pK_b**	**Base Concentration (M)**	**$[H_3O^+]$**	**pH**
CH_3NH_2	4.38×10^{-4}	3.36	0.500	6.85×10^{-13}	12.2
NH_3	1.77×10^{-5}	4.75	0.05	1.08×10^{-11}	11.0
$C_6H_5NH_2$	3.81×10^{-10}	9.42	1.2	4.67×10^{-10}	9.33

LOCK IT IN:
Which of the bases in Task 9 will have the conjugate acid with the lowest pK_a? Explain your reasoning.

$C_6H_5NH_2$ would have the conjugate acid with the lowest pK_a. The lowest pK_a would result from the conjugate base with the largest pK_b. Another way to think of this is that the lowest pK_a would result in the strongest acid and the weaker the conjugate base, the stronger the conjugate acid. $C_6H_5NH_2$ is the weakest base and thus should have the strongest conjugate acid.

CH_3NH_2

$$CH_3NH_2\,(aq) + H_2O\,(\ell) \rightleftharpoons CH_3NH_3^+\,(aq) + OH^-\,(aq)$$

	$[CH_3NH_2]$	$[CH_3NH_2^+]$	$[OH^-]$
Initial	0.500	0.00	0.00
Change	$-x$	$+x$	$+x$
Equilibrium	$0.500-x$	$0.00+x$	$0.00+x$

$$K_b = \frac{[CH_3NH_3^+]\,[OH^-]}{[CH_3NH_2]}$$

$$4.38 \times 10^{-4} = \frac{x^2}{0.500 - x}$$

$$x = \xcancel{-0.0150} \qquad x = 0.0146$$

$$1.00 \times 10^{-14} = [H_3O^+]\,(0.0146)$$

$$[H_3O^+] = 6.85 \times 10^{-13}$$

$$pH = -\log_{10}(6.85 \times 10^{-13}) = 12.2$$

NH_3

$$NH_3\,(aq) + H_2O\,(\ell) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)$$

	$[NH_3]$	$[NH_4^+]$	$[OH^-]$
Initial	0.05	0.00	0.00
Change	$-x$	$+x$	$+x$
Equilibrium	$0.05-x$	$0.00+x$	$0.00+x$

$$K_b = \frac{[NH_4^+]\,[OH^-]}{[NH_3]}$$

$$1.77 \times 10^{-5} = \frac{x^2}{0.05 - x}$$

$$x = \cancel{-0.00095} \qquad x = 0.00093$$

$$1.00 \times 10^{-14} = [H_3O^+]\,(0.00093)$$

$$[H_3O^+] = 1.08 \times 10^{-11}$$

$$pH = -\log_{10}(1.08 \times 10^{-11}) = 11.0$$

$C_6H_6NH_2$

$$C_6H_6NH_2\,(aq) + H_2O\,(\ell) \rightleftharpoons C_6H_6NH_3^+\,(aq) + OH^-\,(aq)$$

	$[C_6H_6NH_2]$	$[C_6H_6NH_3^+]$	$[OH^-]$
Initial	1.2	0.00	0.00
Change	−x	+x	+x
Equilibrium	1.2−x	0.00 +x	0.00 +x

$$K_b = \frac{[C_6H_6NH_3^+]\,[OH^-]}{[C_6H_6NH_2]}$$

$$3.81 \times 10^{-10} = \frac{x^2}{1.2 - x}$$

$$x = -\cancel{2.14 \times 10^{-5}} \qquad x = 2.14 \times 10^{-5}$$

$$1.00 \times 10^{-14} = [H_3O^+]\,(2.14 \times 10^{-5})$$

$$[H_3O^+] = 4.67 \times 10^{-10}$$

$$pH = -\log_{10}(1.08 \times 10^{-11}) = 9.33$$

OBJECTIVE 6

Determine if a solution is acidic, basic, or neutral based on the solute.

Because salts are composed of cations that can sometimes act as acids and anions that are the conjugate bases of acids, salt solutions can be either acidic, basic, or neutral. A few patterns can help us analyze salt solutions:

1. Cations that are small, highly charged metals like iron(III) (Fe^{3+}) and aluminum (Al^{3+}) tend to be weakly acidic. As water molecules arrange themselves around the cation, the positive charge adjacent to the oxygen atom in the water enhances the pull of electrons away from the hydrogen atom. This destabilization of the O—H bond increases the acidity of the hydrogen atom. While this phenomenon is found in solutions of transition metals, it is not the case for alkali or alkaline earth metals. Reference your textbook or lecture notes for more details.

2. Anions that are the conjugate bases of strong acids are so weak that they have no effect on the pH of a solution.

3. Anions that are the conjugate bases of weak acids will act as bases in solution and can raise the pH.

4. In a case where an acidic cation is paired with an anion that acts as a base, whether the solution is acidic, basic, or neutral will depend on the relative strengths of the acid and the base.

OBJECTIVE 6

Determine if a solution is acidic, basic, or neutral based on the solute.

TASK 10: Determine if each salt solution would be acidic, basic, or neutral. Explain your answer in the "Justification" column. Use the sandbox if you are not sure if an anion is the conjugate base of a strong or weak acid.

Solution	Acidic, Basic, or Neutral?	Justification
0.500 M $FeCl_3$	Acidic	A solution containing iron (III) and the conjugate base of a strong acid like HCl will be acidic due to the acidic nature of the Fe^{3+} ion yet pH-neutral nature of the chloride ion (Cl^-).
1.3 M NaCl	Neutral	Neither sodium ions (Na^+) nor chloride ions (Cl^-) have significant effects on the pH of a solution. It will be neutral as a result.
0.72 M KNO_3	Neutral	Neither sodium ions (K^+) nor nitrate ions (NO_3^-) have significant effects on the pH of a solution. It will thus be neutral. Remember that NO_3^- is the conjugate base of the strong acid HNO_3. As such, it is a base so weak that it does not alter the pH of the solution.
1.0 M LiF	Basic	While the lithium ion (Li^+) will not have an effect on the solution, the fluoride ion (F^-) will. Remember that F^- is the conjugate base of a weak acid and will thus have some ability to remove hydrogen ions from water to form HF. The result of this process is the creation of hydroxide ions (OH^-) and a basic solution.
0.74 M AlBr	Acidic	A solution containing aluminum ions (Al^{3+}) and the conjugate base of a strong acid like HBr will be acidic due to the acidic nature of the Al^{3+} ion yet pH-neutral nature of the bromide ion (Br^-).

OBJECTIVE 7

Demonstrate an ability to write a net ionic equation for a neutralization reaction.

Reactions that proceed with at least some components being in an ionic phase can be written in several different ways. Since ions are a very important part of reactions involving acids and bases, their reactions too can be written in multiple ways. An important type of reaction that you completed multiple times during the course of the Acid Strength Game is called a neutralization reaction where an acid and a base react with one another. The usual products of a **neutralization reaction** are water and a soluble salt, but there are occasions where other products are formed.

- **Molecular equations** include all substances written in either their molecular or formula unit form followed by the state of matter of that substance in the reaction. Substances that are liquids or gases receive the designation (*ℓ*) or (*g*). Substances that are solid will receive an (*s*) This designation includes substances that are insoluble in the reaction medium, yet are still present. Soluble compounds receive (aq) after them to indicate that they are dissolved in the solution. Let's use the neutralization reaction between hydroiodic acid (HI) and potassium hydroxide (KOH) and the reaction between acetic acid ($H_2C_2H_3O_2$) and lithium hydroxide (LiOH) as examples.

$$HI\ (aq) + KOH\ (aq) \rightarrow KI\ (aq) + H_2O\ (\ell)$$

$$HC_2H_3O_2\ (aq) + LiOH\ (aq) \rightarrow LiC_2H_3O_2\ (aq) + H_2O\ (\ell)$$

- **Complete ionic equations** separate any substances that fully ionize in solution into their constituent ions, each of which are followed by "(aq)". For example, **if you have a strong acid, the hydrogen ions and anions should be separated.** Notice in the example below that the hydrogen and iodide ions of HI are fully separated. A strong acid like HI fully dissociates in solution.

$$H^+\ (aq) + I^-\ (aq) + K^+\ (aq) + OH^-\ (aq) \rightarrow K^+(aq) + I^-\ (aq) + H_2O\ (\ell)$$

However, any substances that do not fully ionize should maintain the same form as in the molecular equation. This means that **weak acids should not be separated in the complete ionic equation** since most of the acid molecules will not be ionized in solution at any one point in time. Notice in the example below that $HC_2H_3O_2$ (acetic acid) in the reactants has not been separated into its constituent ions. A weak acid like $HC_2H_3O_2$ only partially dissociates in solution.

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

OBJECTIVE 7

Demonstrate an ability to write a net ionic equation for a neutralization reaction.

- **Net ionic equations** are complete ionic equations from which **spectator ions** have been removed. Spectator ions are those which remain in the same aqueous state both before and after the reaction. They do not participate in the reaction and can be identified as appearing the same on both the reactant and product sides of the complete ionic equation. In the example reaction, notice that K+ and I- remain in the same state as both reactants and products. They are spectator ions and are thus removed when creating the net ionic equation.

  The net ionic equation for the reaction between hydroiodic acid and potassium hydroxide is simply the formation of water from hydrogen and hydroxide ions. **This exact equation will be the net ionic equation for any neutralization reaction between a strong acid and strong base.**

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

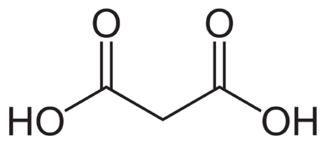
  The net ionic equation for the reaction between acetic acid and lithium hydroxide still maintains the acetate ion ($C_2H_3O_2^-$) since acetic acid was only partially ionized in solution. This will be the case when a weak acid reacts with a strong base.

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

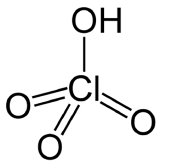
TASK 11: Perform each of the reactions listed below in the sandbox. Afterwards, write the complete ionic and net ionic equation for each in the relevant columns.

Molecular Equation	Complete Ionic Equation	Net Ionic Equation
$HCl\ (aq) + LiOH\ (aq) \rightarrow LiCl(aq) + H_2O\ (\ell)$	$H^+(aq) + Cl^-(aq) + Li^+(aq) + OH^-(aq) \rightarrow Li^+(aq) + Cl^-(aq) + H_2O\ (\ell)$	$H^+\ (aq) + OH^-\ (aq) \rightarrow H_2O\ (\ell)$
$HCN\ (aq) + NaOH\ (aq) \rightarrow NaCN\ (aq) + H_2O(\ell)$	$HCN(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + CN^-(aq) + H_2O\ (\ell)$	$HCN(aq) + OH^-(aq) \rightarrow CN^-(aq) + H_2O\ (\ell)$
$2\ HCl\ (aq) + Mg(OH)_2\ (s) \rightarrow MgCl_2(aq) + 2\ H_2O\ (\ell)$	$2H^+(aq) + 2Cl^-(aq) + Mg(OH)_2\ (s) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq) + 2H_2O\ (\ell)$	$2H^+(aq) + Mg(OH)_2\ (s) \rightarrow Mg^{2+}(aq) + 2H_2O\ (\ell)$

CLOSURE: Below are two molecules that are not found in the sandbox. For each characteristic, determine which value belongs to each molecule using your knowledge of acids and bases. Simply write the correct option in the column underneath the picture of the molecule. Justify your choices in the column to the far right of the table.



Malonic Acid



Perchloric Acid

Property		Options		Justification
Acid Type	Diprotic	**Monoprotic** or **Diprotic**	Monoprotic	At the most basic level, one can see that there are two hydrogen atoms in malonic acid and only one in perchloric acid. Accordingly, perchloric acid can only be monoprotic leaving malonic acid to be diprotic.
Acid Type	Weak Acid	**Strong Acid** or **Weak Acid**	Strong Acid	From what we learned about oxyacids, a highly electronegative atom like chlorine as the center of an acid with four oxygen atoms attached will be quite strong. It is thus designated a strong acid leaving malonic acid to be labeled weak.

First (or Only) pK_a	2.83	**2.83** or **< 1**	<1	The stronger the acid, the smaller the pK_a. Since we have already determined that perchloric acid is a strong acid and malonic acid is a weak acid, the lower pK_a should be given to perchloric acid.
pH of 0.5 M Solution	pH > 1	**pH > 1** or **pH < 1**	pH > 1	A certain concentration of a strong acid will have a lower pH than a weak acid of the same concentration. This is the result of differences in the degree of dissociation for each acid. As such, perchloric acid should be given the lower of the two pH values.