Acids: taste sour

turn blue litmus red

cause carbonates to fizz

cause active metals to “bubble”

react with bases to form water and salts

Bases: feel soapy

taste bitter

turn red litmus blue

cause ammonium salts to “bubble”

react with acids to form water and salts
Reactions with carbonates and ammonium salts are metathesis reactions "with a twist".

\[ \text{CaCO}_3(s) + 2\text{HCl}(aq) \rightarrow \text{CaCl}_2 + \text{H}_2\text{O} + \text{CO}_2 \]

\[ \text{CaCO}_3 + 2\text{H}^+ + 2\text{Cl}^- \rightarrow \]

\[ \text{Ca}^{2+} + 2\text{Cl}^- + \text{H}_2\text{O} + \text{CO}_2 \uparrow \]
\[
\text{NH}_4\text{Cl} \text{ (s)} + \text{NaOH} \text{ (aq)} \rightarrow \text{NH}_4\text{OH} + \text{NaCl} \\
\text{NH}_3 \uparrow + \text{H}_2\text{O}
\]

Arrhenius:

Acids are H\(^+\) donors; bases are OH\(^-\) donors.

\[\text{Acid + Base} \rightarrow \text{Water} + \text{Salt}\]

\[\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{H}_2\text{O} + \text{Na}^+ \text{CH}_3\text{COO}^-\]

\[\text{Na}^+\]

\[\text{hydroxyl group}\]

\[\text{Na}^+\]
End point = equivalence point

1:1 reaction

where pt

\[ \text{mol}_a = \text{mol}_b \]
\[ M_a \cdot V_a = M_b \cdot V_b \]

so

\[ \sqrt{K_a \cdot pH + NaOH \rightarrow H_2O + KNaPh} \]

\[ \text{eqn} \]
\[ \text{mol}_a = \text{mol}_b \]

\[ \text{kg KHP} \left( \frac{1 \text{ mol KHP}}{204.22 \text{ g KHP}} \right) = M_b \cdot V_b \]

known

known
Shortcomings of Arrhenius theory:

1) can distinguish between strong acids and weak acids

\[ \text{HCl} + \text{NaOH} \rightarrow \text{H}_2\text{O} + \text{NaCl} \]
\[ \text{HF} + \text{KOH} \rightarrow \text{H}_2\text{O} + \text{KF} \]
\[ \text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O} \]

but can’t rank weak acids

\[ \text{HF} + \text{OH}^- \rightarrow \text{H}_2\text{O} + \text{F}^- \]

\[ \text{CH}_3\text{COOH} + \text{OH}^- \rightarrow \text{H}_2\text{O} + \text{CH}_3\text{COO}^- \]
2) can’t explain the basicity of compounds that don’t contain hydroxide ions

\[ \text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_4\text{OH} \]

pure fiction

Universal Indicator

\[ \text{pH} \leftrightarrow \text{pK}_a \]

acid \( \uparrow \) neutral \( \downarrow \) base

acid \[ \text{CO}_2 \] base \[ \text{Na}_2\text{CO}_3 \]
In general, a salt that can be viewed as the product of the reaction of a strong acid with a weak base will give aqueous solutions that are acidic.

\[ \text{HCl} + \text{NH}_3 \rightarrow \text{NH}_4\text{Cl} \]

Salt that makes water acidic

In general, a salt that can be viewed as the product of the reaction of a weak acid with a strong base will give aqueous solutions that are basic (alkaline).

\[ \text{H}_2\text{CO}_3 + 2\text{NaOH} \rightarrow 2\text{H}_2\text{O} + \text{Na}_2\text{CO}_3 \]

Salt that makes water basic
$\text{Na}_2\text{Cl} + \text{H}_2\text{O} \rightarrow \text{HCl} + \text{Na}_2\text{OH}$

No!

$\text{HCl} + \text{Na}_2\text{OH} \rightarrow \text{Na}_2\text{Cl} \downarrow$

Salt that doesn't behave as an acid or base.
The reasons for these observations will become more apparent when we discuss the Bronsted-Lowry acid/base theory. For now, we turn our attention to another important class of reactions, oxidation-reduction (redox) reactions.

Like "formal charge", "oxidation numbers (or oxidation states) are based on a set of rules for "bookkeeping" electrons.

The oxidation number is the charge an atom would have if all shared electrons were assigned to the more electronegative atom.

Example: $\text{PO}_4^{3-}$

Formal charges

\[
\begin{align*}
\text{P} & : -1 \\
\text{O} & : 5 - 4 + 0 = +1 \\
\text{O} & : 6 - 1 - \frac{1}{2} = -1
\end{align*}
\]

Oxidation numbers

\[
\begin{align*}
\text{P} & : -1 \\
\text{O} & : 5 - 5 - 0 = 0 \\
\text{O} & : 6 - 2 - 4 = 0
\end{align*}
\]
Ex #10 C -2

\[ \begin{array}{ccc}
\text{O} & \text{P} & \text{O} \\
\text{O} & \text{P} & \text{O} \\
\text{O} & \text{O} & \text{O}
\end{array} \]

\[ 5 \times \]

\[ S + 6 \]

79

can be 0 (whether 10 -

0 = 0)

gets credited with 8e.
Usually, we don’t draw structures for determining oxidation states. We rely on a set of rules that usually permit the assignments to be made without drawing the Dot Diagrams.

The oxidation numbers for a given species must sum to the overall charge.

\[
\begin{align*}
\text{F}_2O_3^- & \quad +9 + 4(-2) = -3 \\
\text{He}, \text{H}_2, \text{I}_3, \text{S}_8 & \\
\text{Na}^+ & \quad 2- \quad \text{Ca}^- \quad \text{Al}^3+
\end{align*}
\]

The oxidation number for an atom in its elemental form is zero.

Group I metals always have an oxidation number of +1 in their compounds.
Group II metals always have an oxidation number of +2 in their compounds.

\[
\begin{array}{c}
\text{Ca} \\
+2 \\
\text{Br} \\
-1
\end{array}
\]

Aluminum always has an oxidation number of +3 in its compounds.

\[
\begin{array}{c}
\text{Al} \\
+3 \\
(\text{OH})_3
\end{array}
\]

Fluorine always has an oxidation number of −1 in its compounds.

\[
\begin{array}{c}
\text{Na} \\
+1 \\
\text{F} \\
-1
\end{array}
\quad \quad \quad
\begin{array}{c}
\text{Ba} \\
+2 \\
\text{F}_2 \\
-1
\end{array}
\]

Hydrogen always has an oxidation number of +1 in its compounds unless it is in a binary compound with an active metal. In such hydrides, hydrogen has an oxidation state of −1.

\[
\begin{array}{c}
\text{H} \\
+1 \\
\text{Cl} \\
-1
\end{array}
\quad \quad \quad
\begin{array}{c}
\text{Na} \\
+1 \\
\text{H} \\
-1
\end{array}
\quad \quad \quad
\begin{array}{c}
\text{Ca} \\
+2 \\
\text{H}_2 \\
-1
\end{array}
\]
Oxygen almost always has an oxidation number of $-2$ in its compounds. However, oxygen has an oxidation number of $-1$ in peroxides. Oxygen will have a positive oxidation number when it is combined with fluorine.

\[
\begin{align*}
\text{H}_2\text{O} & \quad \text{[O-H]}^- \\
+1 & \quad -2
\end{align*}
\]

The other halogens (Cl, Br, and I) have oxidation numbers of $-1$ in binary compounds with metals. When combined with more active non-metals, Cl, Br, and I may have oxidation numbers of $+1$, $+3$, $+5$, and $+7$.

\[
\begin{align*}
\text{Cl}_2\text{F}_2 & \quad \text{I}_2\text{O}_3 \\
+2 & \quad -1 \\
& \quad +5 \quad -2 \\
& \quad +3 \quad -1
\end{align*}
\]

The chalcogens (S, Se, Te) have oxidation numbers of $-2$ in binary compounds with metals. When combined with more active non-metals, S, Se, and Te may have oxidation numbers of $+2$, $+4$, and $+6$.

\[
\begin{align*}
\text{H}_2\text{S} & \quad \text{SO}_4^{2-} \\
+1 & \quad -2 \\
& \quad +6 \quad -2
\end{align*}
\]
\[ \text{H}_2\text{O}_2 \]

\[ \frac{12}{2} \]

\[ 1 \]

\[ \text{As} \text{ is given to } 0 \]

\[ H \]

\[ \text{Shared equally} \]
$\text{OF}_2
\begin{array}{c}
+2 \ -1
\end{array}
\text{Oxygen di fluoride}