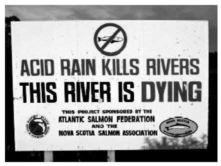
Acids & Bases



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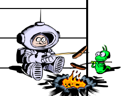
Acids & Bases: Traditional Properties

Property	Acid	Base
Taste	Sour	Bitter
Feel	None	Slippery
Litmus	B➔R	R➔B
Phenolphthalein	Colorless	Magenta
With Carbonate	CO_2 evolution	None
With “active” Metals	H_2 evolution	None
With most metals	None	Water Insoluble



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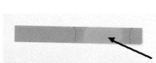
Acids React With Blue Litmus



Blue litmus paper with a drop of acid here

“litmus test” (from lichens … 1300 AD)

Blue ➔ Red in Acid (BRA)

Acids react with carbonate ions:

$2\ H^+_{(aq)} + CO_3^-{}_{(aq)} \rightarrow H_2O_{(l)} + CO_{2(g)}$

Atmospheric $CO_2 + H_2O$ ➔ H_2CO_3

Dissolves Carbonates

A major erosion process

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Activity Series

Metals	Metal Ion	Reactivity
K	K^+	
Ca	Ca^{2+}	reacts with water
Na	Na^+	
Mg	Mg^{2+}	
Al	Al^{3+}	
Zn	Zn^{2+}	
Fe	Fe^{2+}	reacts with acids
Ni	Ni^{2+}	
Sn	Sn^{2+}	
Pb	Pb^{2+}	
H_2	H^+	
Cu	Cu^{2+}	
Hg	Hg^{2+}	
Ag	Ag^+	highly unreactive
Pt	Pt^+	
Au	Au^{3+}	

Acids react with "active" metals:

$2\,H^+_{(aq)} + Zn \rightarrow Zn^{++} + H_{2(g)}$

K, Ca, Na react with water:

$2\,Na + 2\,H_2O \rightarrow 2\,NaOH + H_2$

Mg, Al, Zn, Fe, Ni, Sn, Pb react with acids

$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$

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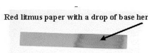
Bases turn Red Litmus ➔ Blue



Bases react with most metal ions:

$2\,OH^-_{(aq)} + M^{2+} \rightarrow M(OH)_{2(s)}$

Most metal hydroxides insoluble in water

Metal Hydroxide Pollution: costly to clean

Bases turn phenolphthalein magenta

Bitter taste plant survival factor

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pH Scale

Measurement of relative acidity

Determined by hydrogen ion concentration

Values commonly range between 0 – 14

- pH < 7 ➔ acidic
- pH = 7 ➔ neutral
- pH > 7 ➔ basic (alkaline)

Measured using

- indicators (pH papers or solutions)
- pH meter

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pH: A Measure of $[H^+]$ (Molar Concentration of H^+)

$[H^+]$	pH
1×10^{-1}	1
1×10^{-2}	2
1×10^{-3}	3
1×10^{-4}	4
1×10^{-5}	5
1×10^{-6}	6
1×10^{-7}	7
1×10^{-8}	8
1×10^{-9}	9
1×10^{-10}	10
1×10^{-11}	11
1×10^{-12}	12
1×10^{-13}	13
1×10^{-14}	14

$[H^+]$ (Acidity) increasing, pH decreasing

$$[H^+] = 1 \times 10^{-pH}$$

$$pH = -\log [H^+]$$

$[H^+]$ (Acidity) decreasing, pH increasing

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pH Scale

Focus of pH scale is the proton (acidity)

Strong acids: pH < 4

Weak acids: pH 4-6

Strong Bases: > pH 11

Weak Bases: pH 8-11

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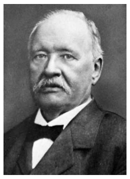


Arrhenius Theory

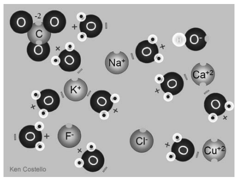
1887 – Svante Arrhenius, Swedish Chemist
- Doctoral Thesis on Electrolytes
- Lowest possible grade

1903 – thesis earned Nobel Prize in Chemistry

Neither water, acids, nor salts conduct
Current only flows by ionization

Acids, special case of ionization

$$HA \rightarrow H^+ + A^-$$



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Arrhenius Theory: Acids

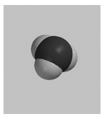
Acid = substance that forms hydrogen ions in water solution

$$HA_{(aq)} \rightarrow H^+_{(aq)} + A^-_{(aq)}$$

H^+ = proton

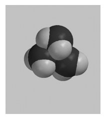
But, individual protons do NOT exist in water:

$$H^+ + H_2O \rightarrow H_3O^+ \text{ (Hydronium Ion)}$$



$\downarrow$ + 2 H_20

$H_7O_3^+$



Arrhenius Acids form *hydronium ions* in solution

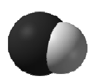
Arrhenius Theory: Bases

Base = substance that forms hydroxide ions (OH^-) in water

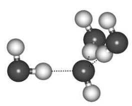
$$MOH_{(aq)} \rightarrow M^+_{(aq)} + OH^-_{(aq)}$$

$$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

$$Ca(OH)_{2\,(aq)} \rightarrow Ca^{+2}_{(aq)} + 2\,OH^-_{(aq)}$$



Arrhenius Bases form *hydroxide ions* in solution



Hydroxide also hydrated ($H_7O_4^-$)

Arrhenius Neutralization Reaction

$$H_3O^+_{(aq)} + OH^-_{(aq)} \rightarrow 2\,H_2O_{(l)}$$

Problems With Arrhenius

Acidic properties depend upon dissociation in aqueous solutions
Fails to predict behavior in non-polar solvents
Bases restricted to the OH^- ion

Problems with Arrhenius Solved in 1923
Johanes Bronsted – Danish Chemist
Martin Lowry – English Chemist

Both simultaneously published … so, name of both on the theory

Allows acids & bases in non-aqueous solutions

Allows bases other than hydroxide

Compound can be either an acid or base dependent on conditions

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Bronsted-Lowry Theory of Acids & Bases

$$AH + B \rightleftarrows BH^{+} + A^{-}$$

Acid = proton donor
Base = proton acceptor (Prime departure from Arrhenius)

Acid-Base reaction = proton transfer

Solvent can be non-aqueous
Bases do not have to have OH^-

water can act as an acid or a base
$HCl + H_2O \rightarrow H_3O^+ + Cl^-$ **(water a base)**
$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$ **(water an acid)**

amphoteric = substance that can act as an acid or as a base

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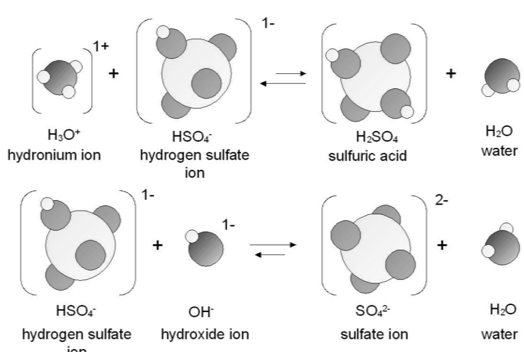
An Amphoteric Ion



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Bronsted-Lowry Neutralization Reactions

$H_3O^+{}_{(aq)} + OH^-{}_{(aq)} \rightarrow 2\,H_2O_{(l)}$

$NH_3 + HCl \rightarrow NH_4^+ + Cl^-$

Arrhenius reactions are also Bronsted-Lowry Acid Base Reactions
But, non-aqueous Bronsted reactions cannot be Arrhenius

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Acid-Base: Conjugate Pairs

$AH + B \rightleftharpoons BH^+ + A^-$

A = Acid (H donor) in forward reaction
B = Base (H acceptor) in forward reaction
BH^+ = Conjugate Acid (H donor in reverse reaction)
A^- = Conjugate Base (H acceptor in reverse reaction)

When asked to find these "conjugate" terms

"Follow the Protons"
'cause
Conjugates differ only by a H^+

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Acid-Base: Conjugate Pairs

$AH + B \rightleftharpoons BH^+ + A^-$

conjugates (base H_2O — acid H_3O^+)

$HCl + H_2O \rightleftharpoons H_3O^+ + Cl^-$
acid, base, acid, base

conjugates (acid HCl — base Cl^-)

HCl	+	H_2O	$\rightleftharpoons$	H_3O^+	+	Cl^-
acid		base		CA		CB

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What are the conjugate acid-base pairs:

A = Acid (H donor) in forward reaction
B = Base (H acceptor) in forward reaction
BH^+ = Conjugate Acid (H donor in reverse reaction)
A^- = Conjugate Base (H acceptor in reverse reaction)

$$NO_2^- + H_2S \leftrightarrow HNO_2 + HS^-$$

A = Acid = H_2S
B = Base = NO_2^-
BH^+ = Conjugate Acid = HNO_2
A^- = Conjugate Base = HS^-

LPT

What are the conjugate acid-base pairs:

A = Acid (H donor) in forward reaction
B = Base (H acceptor) in forward reaction
BH^+ = Conjugate Acid (H donor in reverse reaction)
A^- = Conjugate Base (H acceptor in reverse reaction)

$$HC_4H_5O_3 + PO_4^{3-} \leftrightarrow HPO_4^{2-} + C_4H_5O_3^-$$

A = Acid = $HC_4H_5O_3$
B = Base = PO_4^{3-}
BH^+ = Conjugate Acid = HPO_4^{2-}
A^- = Conjugate Base = $C_4H_5O_3^-$

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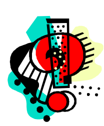
What are the conjugate acid-base pairs:

A = Acid (H donor) in forward reaction
B = Base (H acceptor) in forward reaction
BH^+ = Conjugate Acid (H donor in reverse reaction)
A^- = Conjugate Base (H acceptor in reverse reaction)

$$HSO_4^- + HC_2O_4^- \leftrightarrow SO_4^{2-} + H_2C_2O_4$$

A = Acid = HSO_4^-
B = Base = $HC_2O_4^-$
BH^+ = Conjugate Acid = $H_2C_2O_4$
A^- = Conjugate Base = SO_4^{2-}



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What are the conjugate acid-base pairs:

A = Acid (H donor) in forward reaction
B = Base (H acceptor) in forward reaction
BH^+ = Conjugate Acid (H donor in reverse reaction)
A^- = Conjugate Base (H acceptor in reverse reaction)

$$HNO_2 + CN^- \leftrightarrow NO_2^- + HCN$$

A = Acid = HNO_2
B = Base = CN^-
BH^+ = Conjugate Acid = HCN
A^- = Conjugate Base = NO_2^-

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Fill in the Blanks

Acid	Conjugate Base
HNO_3	
	Br^-
H_2O	
	H_2O
	HPO_4^{2-}
HPO_4^{2-}	
	$C_2H_3O_2^-$

From any compound:

To make conjugate Acid
Add H^+

To make conjugate Base:
Subtract H^+

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Fill in the Blanks

Acid	Conjugate Base
HNO_3	NO_3^-
HBr	Br^-
H_2O	OH^-
H_3O^+	H_2O
$H_2PO_4^-$	HPO_4^{2-}
HPO_4^{2-}	PO_4^{3-}
$C_2H_4O_2$	$C_2H_3O_2^-$

From any compound:

To make conjugate Acid
Add H^+

To make conjugate Base:
Subtract H^+

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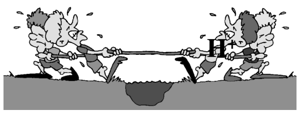
Bronsted-Lowry Theory:

strong acid = excellent proton donor (readily loses H^+)
weak acid = poor proton donor (does not lose H^+ easily)

strong base = very good proton acceptor (readily gains H^+)
weak base = poor proton acceptor (does not gain H^+ easily)
The stronger the attraction for H^+, the stronger the base

"Weak" or "Strong" is about "H^+ attraction"
Strong Acid has Weak Conjugate Base
Weak Acid has Strong Conjugate Base
Strong Base has Weak Conjugate Acid
Weak Base has Strong Conjugate Acid



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Relative Strengths of Acids & Bases

"strong" acid or base: 100 % completely ionized

$HCl \rightarrow H^+ + Cl^-$
$HNO_3 \rightarrow H^+ + NO_3^-$

"weak" acid or base: < 100 % ionized, partially ionized

$HC_2H_3O_2 \leftrightarrow H^+ + C_2H_3O_2^-$
$HF \leftrightarrow H^+ + F^-$

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pH & pK_a

pH = negative logarithm of $[H^+]$
K_a = a measure of acidity based on dissociation:
pK_a = negative logarithm of K_a

$HA \rightleftharpoons H^+ + A^-$

K_a = measure of acidity

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

Lots of [HA] lowers K_a (acidity)

Strong Acid:

$HA \rightleftharpoons H^+ + A^-$

Examples: HCl, HNO_3

Weak Acid:

$HA \rightleftharpoons H^+ + A^-$

Example: Acetic Acid

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Auto-ionization (Self-Protolysis) of Water

$H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$

At 25 °C:
[H^+] = 1 x 10^{-7} M
[OH^-] = 1 x 10^{-7} M

[ ] means Molar (moles / L) Concentration

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Ion-Product Constant for Water (K_w)

K_w = Product of [H_3O^+] multiplied by [OH^-]

$[H_3O^+][OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$

$K_w = 1.0 \times 10^{-14}$

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pH: A Measure of [H^+] (Molar Concentration of H^+)

[H^+] > [OH^-] ➔ acidic solution
[OH^-] > [H^+] ➔ basic solution
[H^+] = [OH^-] ➔ neutral solution

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

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Calculate $[H^+]$ for an aqueous solution at 25°C with a $[OH^-]$ of 1.0×10^{-5} M; Is this solution is acidic, neutral, or basic?

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

$[H^+][1.0 \times 10^{-5}] = K_w = 1.0 \times 10^{-14}$

$[H^+][1.0 \times 10^{-5}] = 1.0 \times 10^{-14}$

$$[H^+] = \frac{1.0 \times 10^{-14}}{[1.0 \times 10^{-5}]}$$

$[H^+] = 1.0 \times 10^{-9}$

$[OH^-] > [H^+]$ ➔ basic solution

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Calculate pH of a solution with $[H^+]$ of 1.0×10^{-5} M.

$pH = -\log [H^+]$

$pH = -\log (1.0 \times 10^{-5})$

$pH = 5.00$

Number of sig figs in the original number:
number of ***decimal places*** in the number after taking the log

Calculate pH of a solution with $[H^+]$ of 4.8×10^{-3} M.

$pH = -\log [H^+]$

$pH = -\log (4.8 \times 10^{-3})$

$pH = 2.32$

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What is the $[OH^-]$ of a solution with a pOH of 7.43?

"p N" means – log of [N]

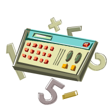
$pOH = -\log [OH^-]$

$7.43 = -\log [OH^-]$ (2 decimal places)

$-7.43 = \log [OH^-]$

$[OH^-] = \text{anti-log} - 7.43$

$[OH^-] = 3.71535 \times 10^{-8}$ ➔ 3.7×10^{-8} (2 sig figs)



Calculator function depends on keypad:
Anti-log = inverse log = inv $=10^x = y^x$
Must try your calculator to validate keystrokes

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pH Scale

$[H^+][OH^-] = Kw = 1x10^{-14}$

$pH + pOH = 14.00$

The pH of a solution is 8.23; what is the pOH?

$pH + pOH = 14.00$

$pOH = 14 – pH$

$pOH = 14 – 8.23$

$pOH = 5.77$

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Calculate the pH of a 0.010M solution of HCl.

Strong Acid: $HCl \rightleftharpoons H^+ + Cl^-$

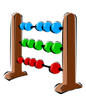
Strong Acid ➔ Assume M = [H⁺]

$[H^+] = 0.010M = 1.0 \times 10^{-2} M$

$pH = -\log [H^+]$

$pH = -\log (1.0 \times 10^{-2})$

$pH = 2.00$



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Relative Acid and Base Strengths

	Acid	K_a (at 25°C)	Conjugate Base	
Strongest Acid	HI	10^{11}	I^-	Weakest Base
↓	HBr	10^{9}	Br^-	↑
	$HClO_4$	10^{7}	ClO_4^-	
	HCl	10^{7}	Cl^-	
	H_2SO_4	10^{2}	HSO_4^-	
	HNO_3	20	NO_3^-	
	H_3O^+	1	H_2O	
	H_2SO_3	1.5×10^{-2}	HSO_3^-	
	HSO_4^-	1.2×10^{-2}	SO_4^{2-}	
	H_3PO_4	7.5×10^{-3}	$H_2PO_4^-$	
	HF	7.2×10^{-4}	F^-	
	HNO_2	4.0×10^{-4}	NO_2^-	
	HCO_2H	1.8×10^{-4}	CO_2H^-	
	$HC_2H_3O_2$	1.8×10^{-5}	$C_2H_3O_2^-$	
	H_2CO_3	4.3×10^{-7}	HCO_3^-	
	HSO_3^-	1.0×10^{-7}	SO_3^{2-}	
	H_2S	9.1×10^{-8}	HS^-	
	HClO	3.5×10^{-8}	ClO^-	
	HBrO	2.0×10^{-9}	BrO^-	
	HCN	6.2×10^{-10}	CN^-	
	NH_4^+	5.6×10^{-10}	NH_3	
	HCO_3^-	4.8×10^{-11}	CO_3^{2-}	
	H_2O_2	2.4×10^{-12}	HO_2^-	
Weakest Acid	H_2O	1.0×10^{-14}	OH^-	Strongest Base

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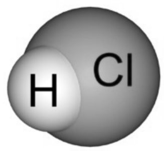
So, I gave that base a proton



Bases love protons

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