CHEM 102 B
Lecture 3: Acids and Bases

At a glance: Kw, pH, ICE calculations, conjugate pairs, and Lewis acids.

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

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



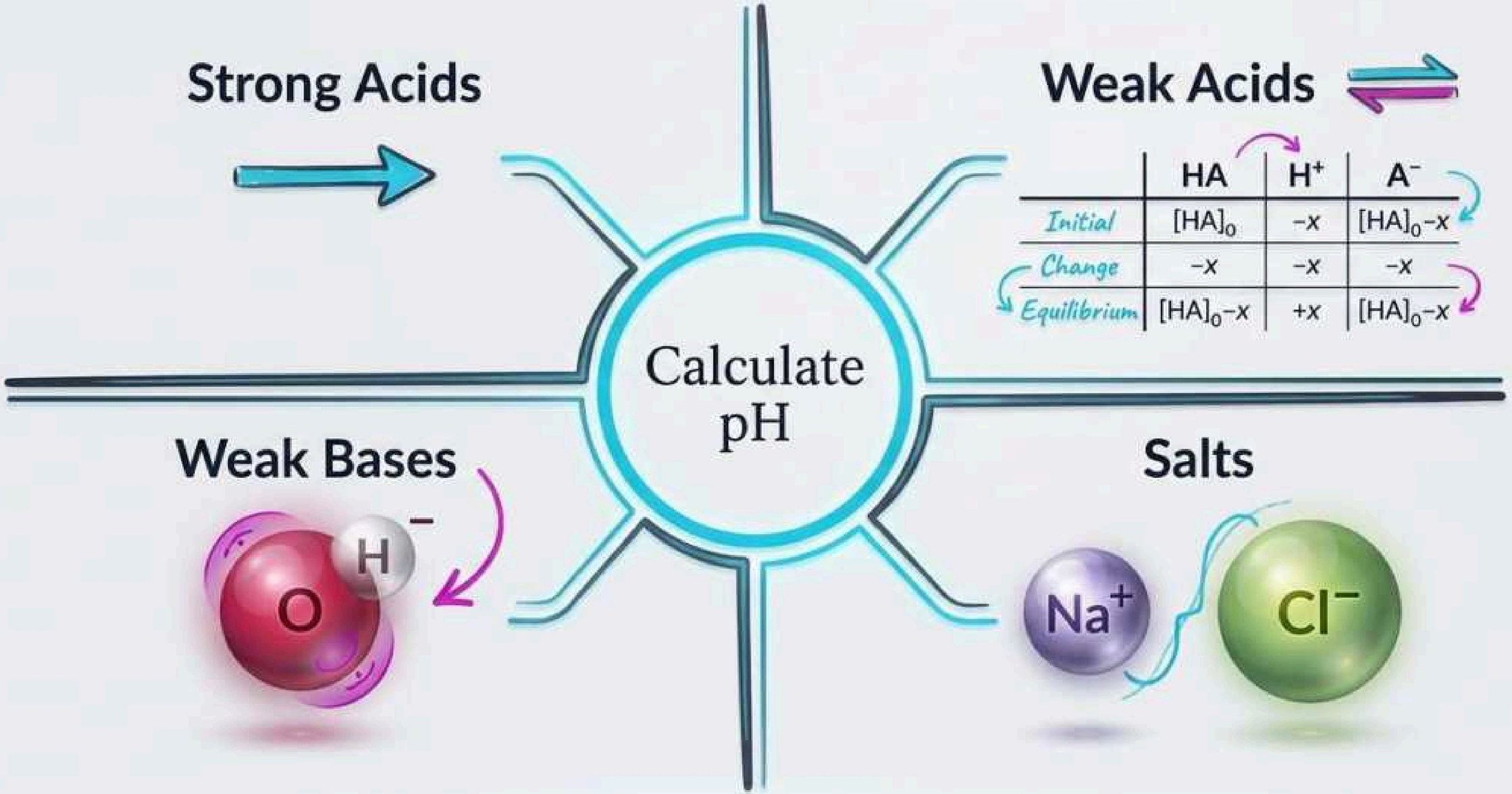
$$\mathrm{pH} = -\log[H^+]$$

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

We now have the mathematical tools to calculate pH properly

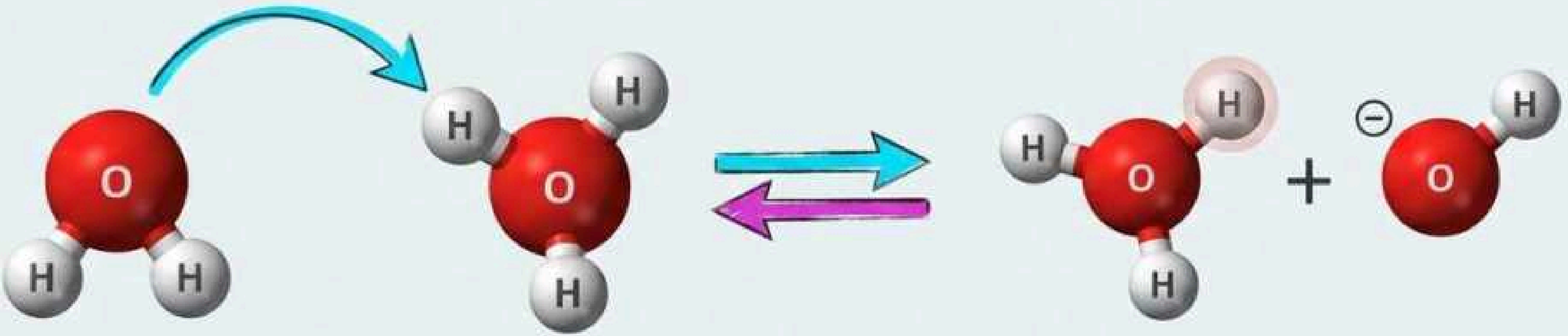
¶ 3.2

Chemistry One introduced acids, bases, and pH in a few paragraphs. Now, we treat weak acids as true equilibria requiring ICE tables.



Autoionisation:

Water reacts with itself



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

At 25°C. Only a tiny fraction of molecules do this.

The strict balance of pH and pOH

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



pH + pOH = 14

If you know one, you know the other: A pH of 11 means a pOH of 3.

Strong acids and bases ionise completely

$0.010\ M\ \mathrm{HCl} \xrightarrow{100\%} [\mathrm{H_3O^+}] = 0.010\ \mathrm{M}$

pH = 2

$0.0010\ M\ \mathrm{NaOH} \xrightarrow{100\%} [\mathrm{OH^-}] = 0.0010\ \mathrm{M}$

pOH = 3
→ pH = 11

Typical examples: Hydrochloric, nitric, and sulfuric acids.

Weak acids ionise only partly (Ka)

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]\cancel{[H_2O]}}$$

Water is left out because it is the solvent.

Translating strength: K_a versus pK_a

¶ 3.7

Acetic Acid (Vinegar)	$K_a = 1.8 \times 10^{-5}$	$pK_a = 4.74$
Hydrofluoric Acid	$K_a \approx 6 \times 10^{-4}$	(Stronger than acetic)
Hydrochloric Acid	$K_a = \infty$ (Enormous)	(Gap is many orders of magnitude)

Stronger Acid
(Smaller pK_a)

Weaker Acid
(Larger pK_a)

pK_a is simply -log(K_a).
A small K_a means a weak acid.

Building the ICE Table for 0.10 M Acetic Acid

	[HA]	[H_3O^+]	[A^-]
I (Initial)	0.10	~ 0	0
C (Change)	-x	+x	+x
E (Equilibrium)	0.10 - x	x	x

$$K_a = 1.8 \times 10^{-5} = \frac{x^2}{0.10 - x}$$

The approximation math for weak acids

$$1.8 \times 10^{-5} = \frac{x^2}{0.10 - x}$$

$$1.8 \times 10^{-5} = \frac{x^2}{0.10 - \cancel{x}}$$

The acid is so weak that x is tiny compared to 0.10. Approximate denominator as 0.10.

Simplify: $x^2 = 1.8 \times 10^{-6}$

Solve: $x \approx 1.3 \times 10^{-3}\, M$

x is the $[H_3O^+]$. Taking the negative logarithm gives a $pH \approx 2.88$.

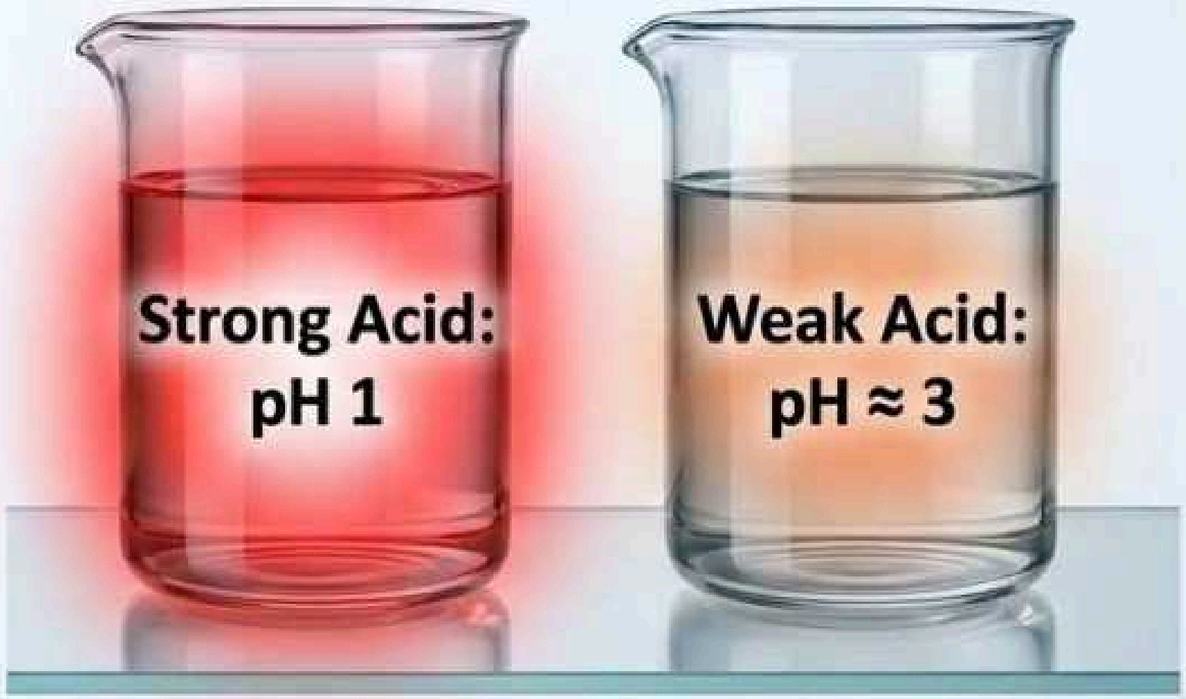
Checking the assumption and seeing the contrast

The 5% Rule Check

$$\left(\frac{1.3 \times 10^{-3}}{0.10}\right) \times 100 = 1.3\%$$

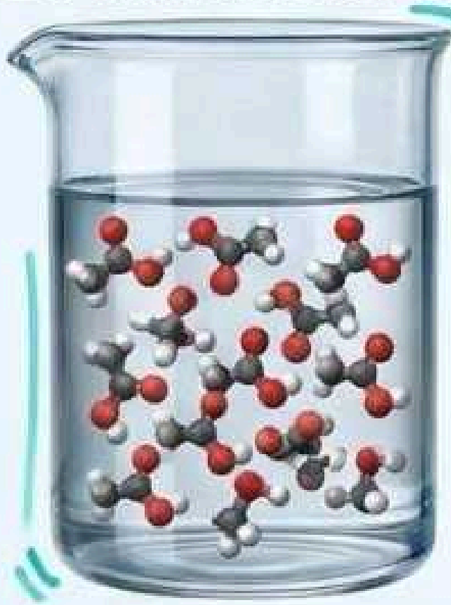
(Note: If over 5%, solve the quadratic)

Identical concentrations: 0.10 M



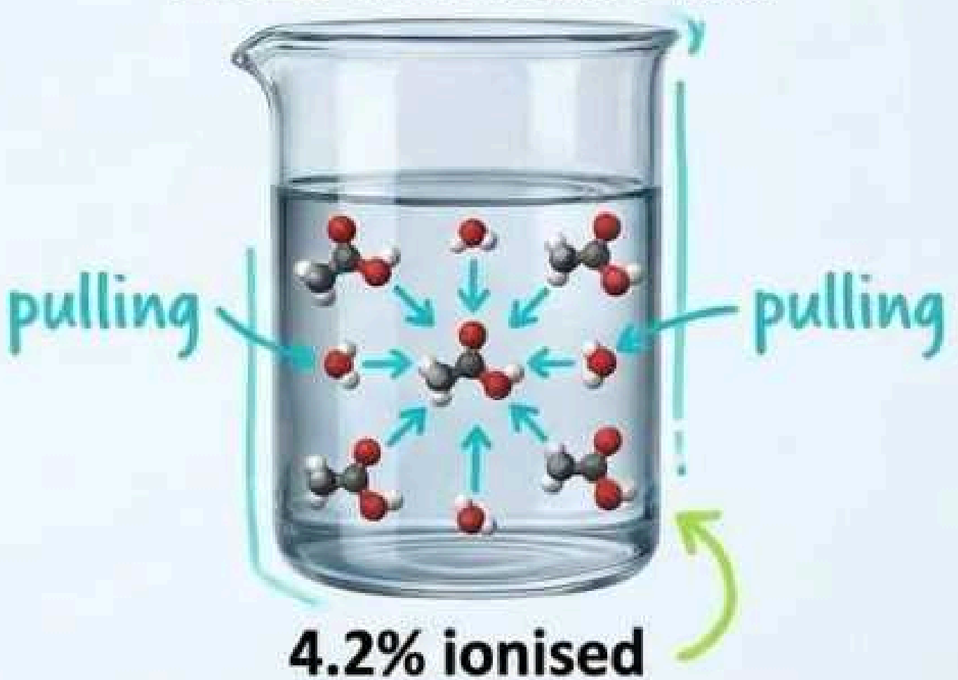
Le Chatelier's Principle: Dilution increases ionisation

0.10 M Acetic Acid



1.3% ionised

0.010 M Acetic Acid



4.2% ionised

A **surprising fact:** A weak acid ionises more, as a fraction, the more dilute it is. Dilution pushes the equilibrium toward more ions.

Weak bases behave mathematically like weak acids

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

$$K_b = 1.8 \times 10^{-5}$$

Same ICE math $\rightarrow [OH^-] = 1.3 \times 10^{-3}\ M$

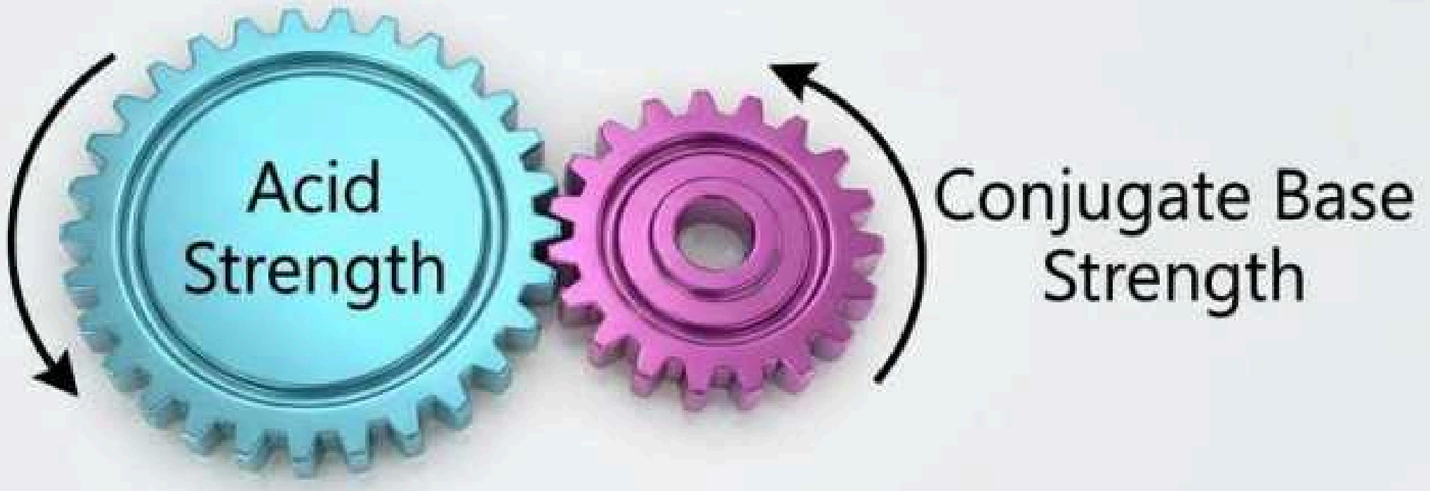
$\rightarrow pOH = 2.88$

$\rightarrow pH = 11.12$

A **weak base** takes a hydrogen ion from water, producing hydroxide.

Conjugate pairs: The inverse strength relationship

Acid – H^+ ⟶ Conjugate Base



$$K_a \times K_b = K_w$$

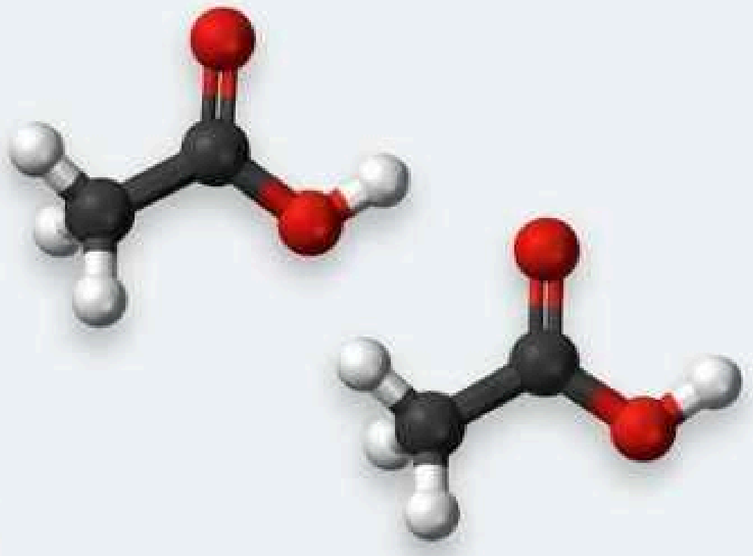
- Strong acid = very weak conjugate base.
- Weak acid = conjugate base of useful strength.

You can always get one constant from the other.

Calculating K_b for a conjugate base

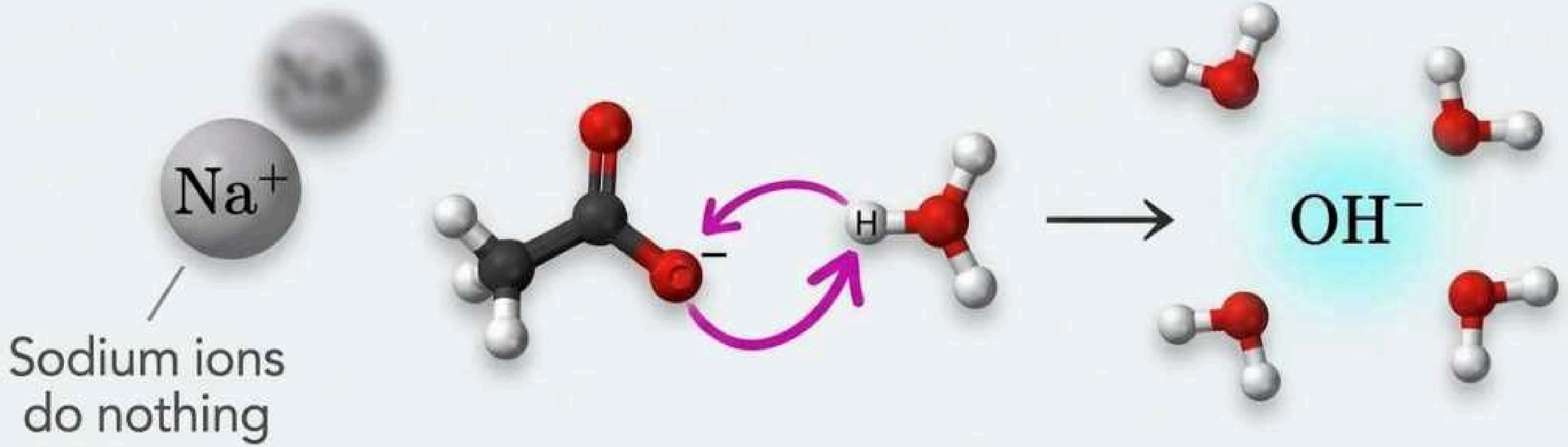
$$K_b = \frac{K_w}{K_a}$$

$$K_b = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$



!!! Acetate is a very weak base, but it is not nothing. That fact explains why some salts are not neutral.

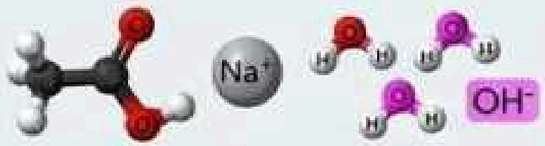
Salts of weak acids create basic solutions



$$[\mathrm{OH^-}] = \sqrt{K_\mathrm{b} \times C} \approx 7.5 \times 10^{-6}\ \mathrm{M}$$

$$\mathrm{pOH} = 5.13 \rightarrow \boxed{\mathbf{pH = 8.87}}$$

The Salt Hydrolysis Diagnostic Matrix

Strong Acid + Strong Base →	Neutral (e.g., *NaCl*) 
Weak Acid + Strong Base →	Basic (e.g., *Sodium Acetate*) 
Strong Acid + Weak Base →	Acidic (e.g., NH_4Cl) 

(Note: Ammonium chloride is acidic because NH_4^+ is the conjugate acid of ammonia and donates an H^+ to water).

Diagnostic Rule

To predict a salt's effect, ask: Which ion has a conjugate in a weak acid or weak base?
That ion reacts with water. The other does not.

Polyprotic acids give up protons in stages

Phosphoric acid (H_3PO_4)

Step 1 (K_{a1}):
7.5×10^{-3}

Sets the pH

Step 2 (K_{a2}):
6.2×10^{-8}
(~100,000x weaker)

Step 3 (K_{a3}):
$\sim 4 \times 10^{-13}$
(~100,000x weaker)

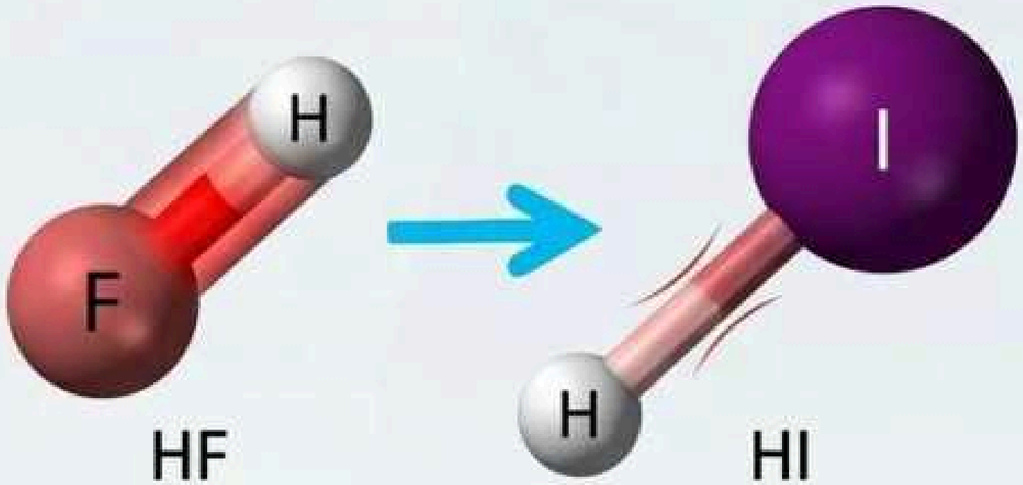
Each step has its own K_a.

Because each step is roughly 100,000 times weaker, the later steps contribute little.

For most purposes, the first ionisation sets the pH.

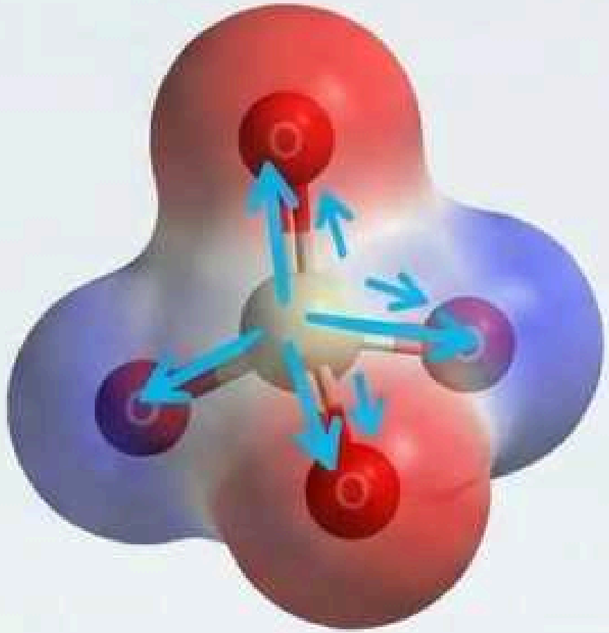
What physical traits make an acid stronger?

1. **Bond strength** must be **weak** enough to break.



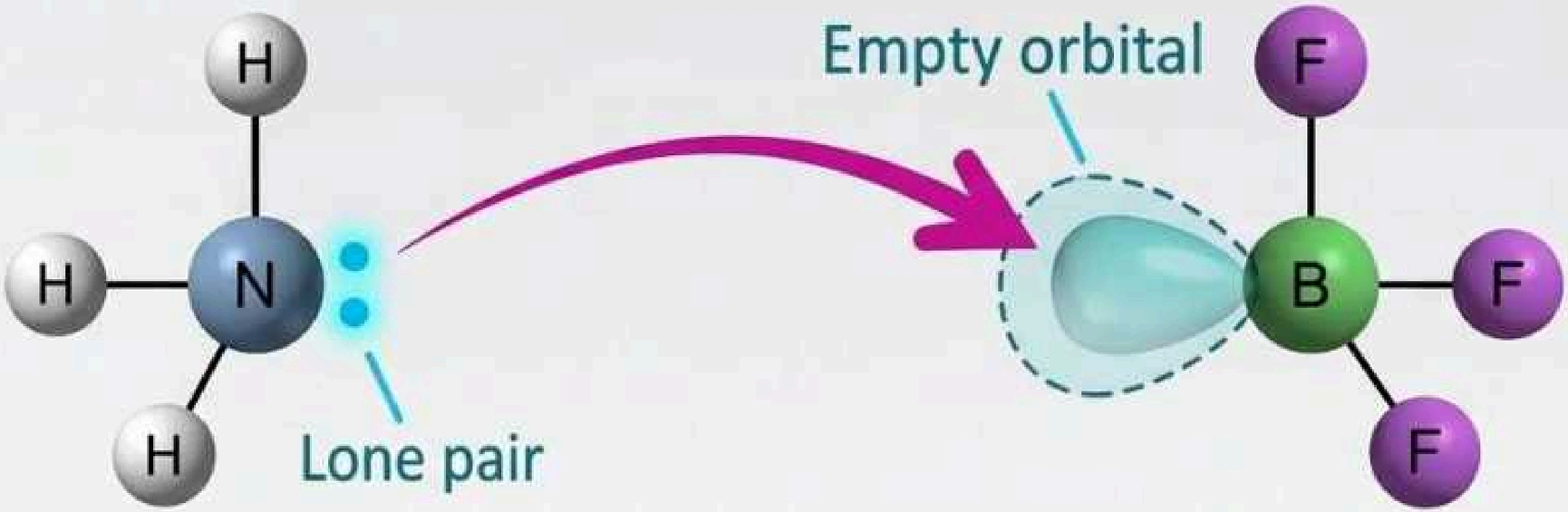
Weaker bond to hydrogen = stronger acid (HI > HF).

2. **Conjugate base** must be **stable.**



Stable conjugate base (delocalized charge) = stronger acid.

The Lewis Definition: Beyond Hydrogen



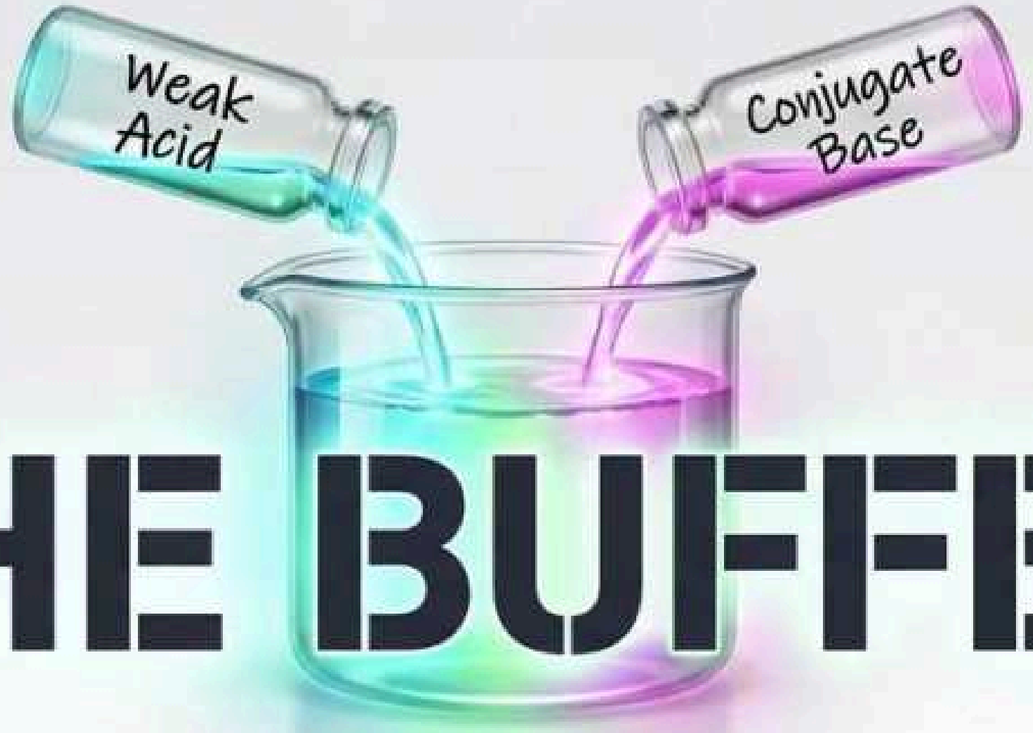
Lewis Base = Electron pair donor

Lewis Acid = Electron pair acceptor

Gilbert Lewis offered a broader definition. Every Brønsted base is a Lewis base, but Lewis acids include molecules with no hydrogen at all, like BF_3 or metal ions bonding to ammonia.

You can now find the pH of almost any simple solution

✓ Strong Acids ✓ Weak Acids (ICE) ✓ Weak Bases ✓ Conjugate Salts ✓ Polyprotic



THE BUFFER

Coming up next: Mixing a weak acid and its conjugate base together creates one of the most important mixtures in chemistry and biology.