8.7 Acid Base Properties of Pure Water

Autoionization of Water

$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+ + OH^-$

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

↳ $K_w = 1 \times 10^{-14}$ @ 25°C

@ 25°C

$[H_3O^+][OH^-] = 1 \times 10^{-14}$

$[H_3O^+] = 1 \times 10^{-7}$

$[OH^-] = 1 \times 10^{-7}$

NEUTRAL: $[H_3O^+] = [OH^-]$

ACIDIC: $[H_3O^+] > [OH^-]$

BASIC: $[H_3O^+] < [OH^-]$

8.7 Acid Base Properties of Pure Water

Curiosity!

The autoionization of water is an endothermic process.
$H_2O(l) + H_2O(l) \Leftrightarrow H_3O^+ + OH^-$

Thus as the temperature increases then – the $[H_3O^+]$ should – 

a) Decrease **b) Increase ✓** **c) Remain the same**

$H_2O(l) + H_2O(l) + \text{'heat'} \Leftrightarrow H_3O^+ + OH^-$

- Increasing T the equivalent of adding reactant
- Shift towards products ... more H_3O^+ produced.

8.7 Acid Base Properties of Pure Water

Curiosity!

With the $[H_3O^+]$ increasing with increasing temperature this must mean that as the temperature of water increases the water – 

a) becomes acidic **b) becomes basic**
c) remain neutral ✓

$H_2O(l) + H_2O(l) + \text{'heat'} \Longleftrightarrow H_3O^+ + OH^-$

↳ As per previous slide there is a shift towards products
BUT
to each H_3O^+ produced there is also an OH^-

8.7 Acid Base Properties of Pure Water

Example I

An aqueous solution has a hydronium ion, H_3O^+, concentration of 1×10^{-11}M @ 25°C. This solution is –

a) acidic b) basic ✓ c) neutral

$K_w = 1 \times 10^{-14}$ @ 25°C

$$[H_3O^+][OH^-] = 1 \times 10^{-14}$$
$$(1 \times 10^{-11})[OH^-] = 1 \times 10^{-14}$$

$$[OH^-] = \frac{1 \times 10^{-14}}{1 \times 10^{-11}} = 1 \times 10^{-3}$$

$$[OH^-] > [H_3O^+]$$

8.8 What are pH and pOH?

$$pH = -\log_{10}[H_3O^+] \quad ; \quad pOH = -\log_{10}[OH^-]$$

The following is only for information purposes... the final formula all you need.

$$[H_3O^+][OH^-] = 1\times10^{-14} \text{ @ } 25^\circ C$$

Take $\log_{10}$ of both sides :

$$\log_{10}([H_3O^+][OH^-]) = \log_{10} 1\times10^{-14}$$

$$\log_{10}[H_3O^+] + \log_{10}[OH^-] = -14$$

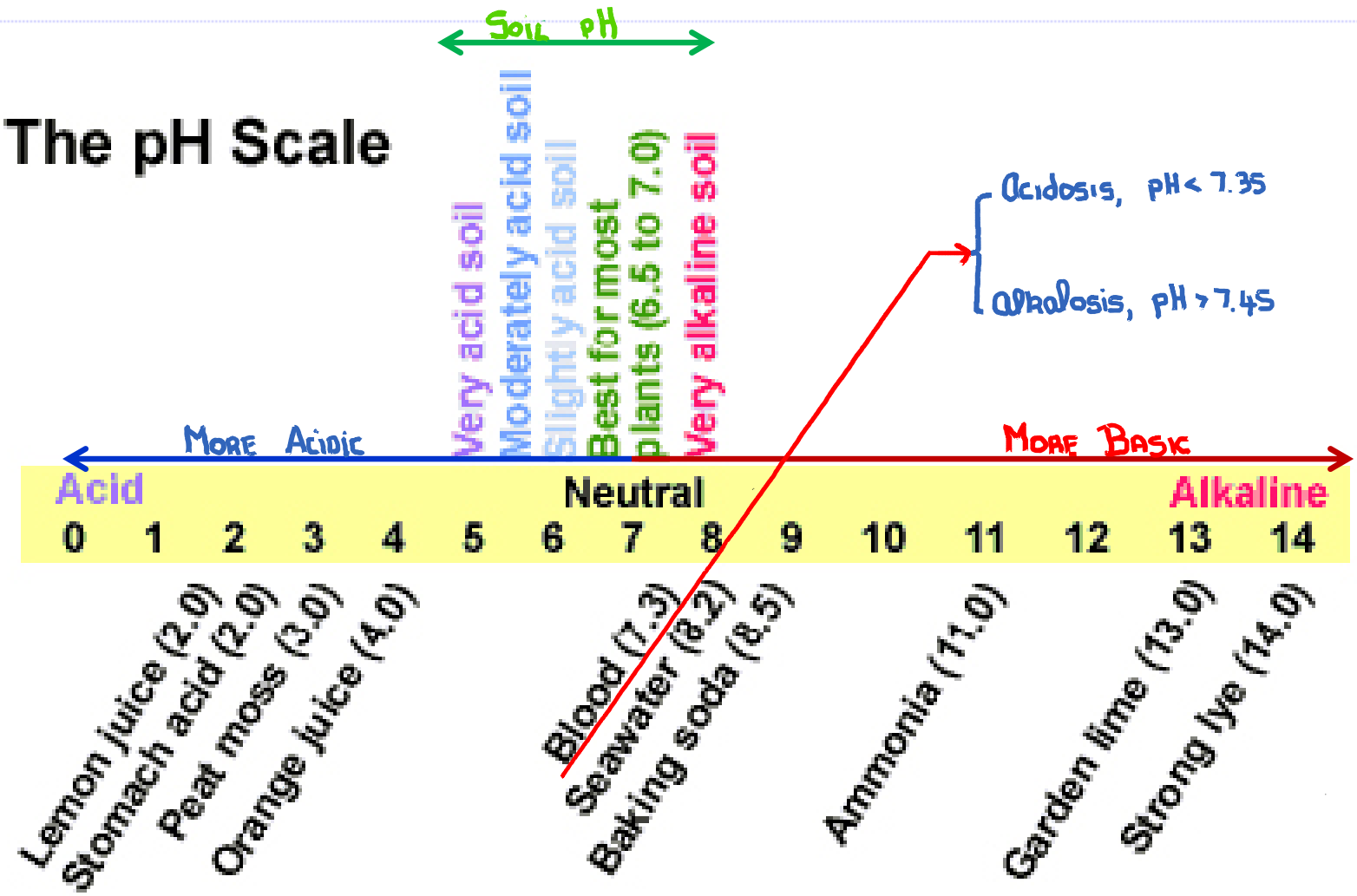
Multiply both sides by -1 :

$$\underbrace{-\log_{10}[H_3O^+]}_{pH} \underbrace{-\log_{10}[OH^-]}_{pOH} = 14$$

$$pH + pOH = 14 \text{ @ } 25^\circ C$$

8.8 What are pH and pOH?

pH – Acidity and Basicity



9.8 What are pH and pOH?

pH – Acidity and Basicity

Plant Preferences for pH			
Very acid	Moderately acid	Slightly acid	Very alkaline
5.0 - 5.8	5.5 - 6.8	6.0 - 6.8	7.0 - 8.0
azalea	bean	asparagus	acacia
blueberry	begonia	beet	bottlebrush
celeriac	Brussels sprouts	bok choy	cabbage
chickory	calla	broccoli	cauliflower
crabapple	camellia	gooseberry	celery
cranberry	carrot	grape	Chinese cabbage
eggplant	collard greens	kale	cucumber
endive	corn	kohlrabi	date palms
heathers	fuchsia	lettuce	dusty miller
huckleberry	garlic	mustard	eucalyptus
hydrangea	lima bean	muskmelon	geranium
Irish potato	parsley	oats	oleander
lily	pea	okra	olive
lupine	peppers	onion	periwinkle
oak	pumpkin	pansy	pinks
raspberry	radish	peach	pomegranate
rhododendron	rutabaga	peanut	salt cedar
rhubarb	soybean	pear	tamarisk
shallot	squash	peony	thyme
sorrel	sunflower	rice	
spinach beet	tomato	spinach	
spruce	turnip	Swiss chard	
wild strawberry	viola		
sweet potato			
watermelon			
white birch			

8.8 What are pH and pOH?

pH – Acidity and Basicity – Example I

An aqueous solution has an $[OH^-] = 1 \times 10^{-5}$ – the pH of this solution is:

$$pOH = -\log_{10}(1 \times 10^{-5})$$
$$= 5$$

$$pH + pOH = 14$$
$$pH = 14 - 5 = 9$$

8.8 What are pH and pOH

pH – Acidity and Basicity – Example II

a) A 0.15M aqueous solution of an acid HA has a measured pH equal to 0.82
b) A 0.45M aqueous solution of an acid HB has a measured pH equal to 0.69 ✓
c) Tom, I have no idea.

Which solution is more acidic? 

The more acidic solution ... the one with the smallest pH.

8.8 What are pH and pOH

pH – Acidity and Basicity – Example III

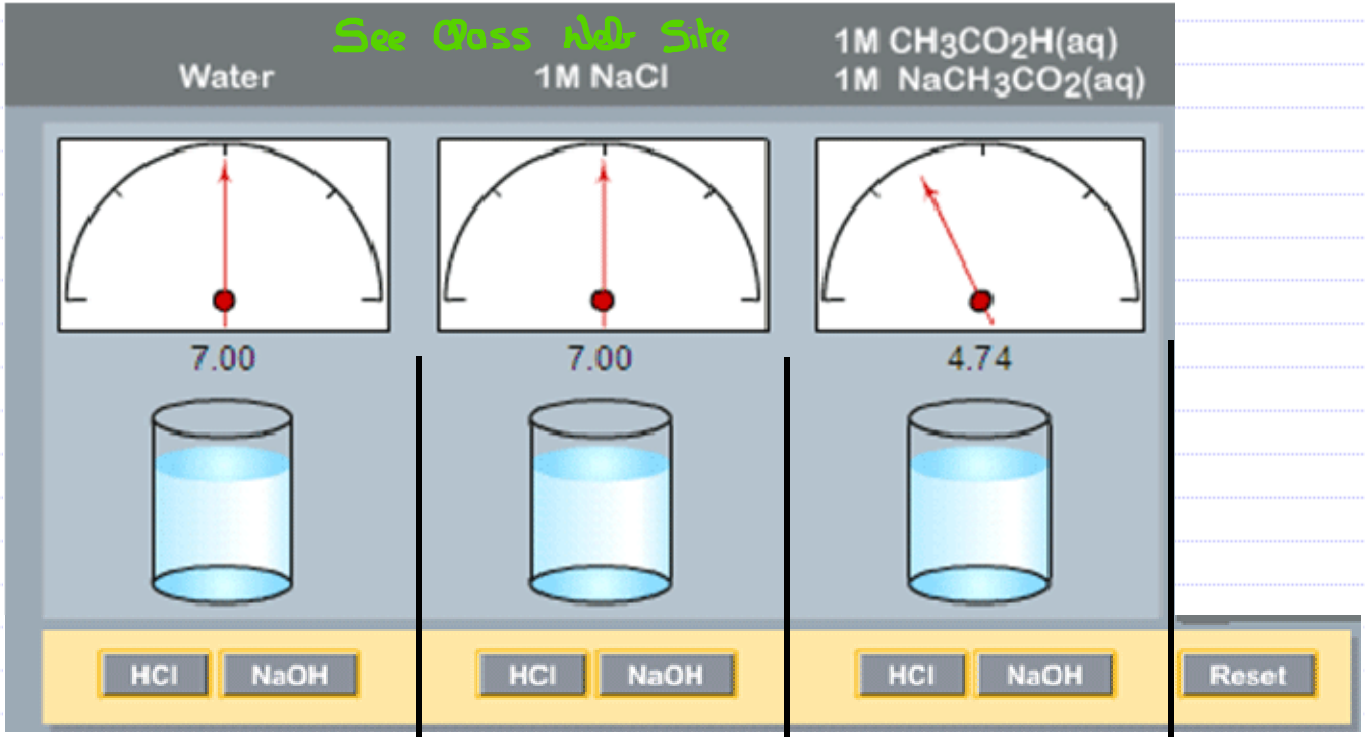
a) A 0.15M aqueous solution of an acid HA has a measured pH equal to 0.82 ✓
b) A 0.45M aqueous solution of an acid HB has a measured pH equal to 0.69
c) Tom, I have no idea.

Which is the stronger acid? 

HA) $pH = -\log_{10}[H_3O^+]$
$= -\log_{10}(0.15) = 0.82$
↳ Expected pH if HA is strong ... goes 100% to ions.

HB) $pH = -\log_{10}[H_3O^+]$
$= -\log_{10}(0.45) = 0.35$
↳ Expected pH if HB is strong ... goes 100% to ions
Actual pH = 0.69, much less acidic than expected ...
Thus HB is a weaker acid than HA.

8.10 What Are Buffers?



pH	7.00	7.00	4.74	Initial
pH	1.04	1.04	4.65	Added H_3O^+
pH	12.96	12.96	4.83	Added OH^-

Large pH changes

Small pH changes!
Why does this solution resist large fluctuations in pH