Purpose What is one way to tell how dangerous an acid or base is?

WARMUP: Try to use a calculator to calculate

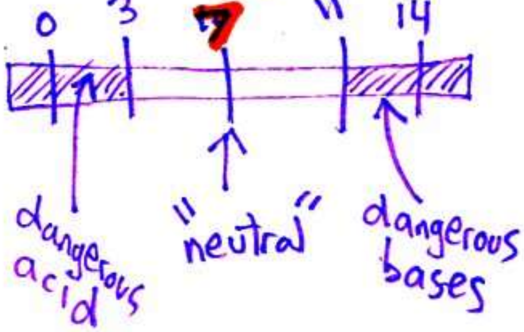
$$-\log\left(5.6\times10^{-8}\right)$$

PUNCH THIS

$\boxed{5}$ $\boxed{.}$ $\boxed{6}$ $\boxed{EXP}$ $\boxed{(-)}$ $\boxed{8}$ $\boxed{LOG}$

ANSWER: -7.25

#2 The pH scale.



#3 ALL WATER SOLUTIONS HAVE H^+ AND OH^-

FORMULA TO MEMORIZE:

$$1 \times 10^{-14} = \begin{pmatrix} \text{concentration} \\ \text{of} \\ OH^- \end{pmatrix} \begin{pmatrix} \text{concentration} \\ \text{of} \\ H^+ \end{pmatrix}$$

FORMULA TO MEMORIZE:

$$pH = -\left(\log [H^+]\right)$$

(use $pH = -\log[H^+]$ formula)

TRY THESE:

0.005	?
0.050	?
0.0005	?
?	6.5
?	2.5
?	0.5

ANSWERS TO "TRY THESE"

$[H^+]$	pH
0.005	2.3
0.050	1.3
0.0005	3.3
3.16×10^{-7}	6.5
3.16×10^{-3}	2.5
0.316	0.5

Table 20.2

Relationship among $[H^+]$, $[OH^-]$, and pH

	$[H^+]$ (mol/L)	$[OH^-]$ (mol/L)	pH	Aqueous system
Increasing acidity ↑	1×10^{0}	1×10^{-14}	0.0	1*M* HCl
	1×10^{-1}	1×10^{-13}	1.0	0.1*M* HCl; Gastric juice
	1×10^{-2}	1×10^{-12}	2.0	Lemon juice
	1×10^{-3}	1×10^{-11}	3.0	
	1×10^{-4}	1×10^{-10}	4.0	Tomato juice
	1×10^{-5}	1×10^{-9}	5.0	Black coffee
	1×10^{-6}	1×10^{-8}	6.0	Milk
Neutral	1×10^{-7}	1×10^{-7}	7.0	Pure water; Blood
	1×10^{-8}	1×10^{-6}	8.0	Sodium hydrogen carbonate, sea water
	1×10^{-9}	1×10^{-5}	9.0	
	1×10^{-10}	1×10^{-4}	10.0	Milk of magnesia
	1×10^{-11}	1×10^{-3}	11.0	Household ammonia
	1×10^{-12}	1×10^{-2}	12.0	Washing soda
	1×10^{-13}	1×10^{-1}	13.0	0.1*M* NaOH
Increasing basicity ↓	1×10^{-14}	1×10^{0}	14.0	1*M* NaOH

In a definition similar to that of pH, the pOH of a solution equals the negative logarithm of the hydroxide-ion concentration.

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

A neutral solution has a pOH of 7. A solution with a pOH less than 7 is basic. A solution with a pOH greater than 7 is acidic. A simple relationship between pH and pOH makes it easy to find either one when the other is known.

$$pH + pOH = 14$$
$$pH = 14 - pOH$$
$$pOH = 14 - pH$$

For pH calculations, you should express the hydrogen-ion concentration in scientific notation. For example, a hydrogen-ion concentration of $0.0010M$, rewritten as $1.0 \times 10^{-3}M$ in scientific notation, has two significant figures. The pH of this solution is 3.00, with the two numbers to the right of the decimal point representing the two significant figures in the concentration. A solution with a pH of 3.00 is acidic, as shown in **Figure 20.6.** How many significant figures are indicated in a pH of 7.61?

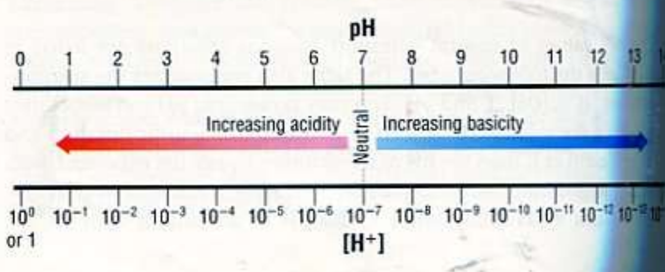


Figure 20.6
The pH scale shows the relationship between pH and the hydrogen-ion concentration. Notice that acids have lower pHs than bases.

Acid–Base Indicators An indicator (In) is an acid or a base that undergoes dissociation in a known pH range. An indicator is a valuable tool for measuring pH because its acid form and base form have different colors in solution. The following generalized equation represents the dissociation of an indicator (HIn)

$$\underset{\text{Acid form}}{\mathrm{HIn}(aq)} \underset{\mathrm{H^+}}{\overset{\mathrm{OH^-}}{\rightleftharpoons}} \mathrm{H^+}(aq) + \underset{\text{Base form}}{\mathrm{In^-}(aq)}$$

The acid form dominates the dissociation equilibrium at low pH (high $[H^+]$) and the base form dominates the equilibrium at high pH (high $[OH^-]$). For each indicator, the change from dominating acid form to dominating base form occurs in a narrow range of approximately two pH units. Within this narrow range, the color of the solution is a mixture of the colors of the acid and the base forms. Knowing the pH range over which this color change occurs can give you a rough estimate of the pH of a solution. At all pH values below this range, you would see only the color of the acid form. At all pH values above this range, you would observe only the color of the base form. You could eventually zero in on a more precise estimate of the pH of the solution by repeating the experiment with indicators that have different pH ranges for their color changes. Many different indicators are needed to span the entire pH spectrum. **Figure 20.8** shows the pH ranges of some commonly used indicators.

Figure 20.8

Each indicator changes color at a different pH. Which indicator would you choose to show that a reaction solution has changed from pH 3 to pH 4?