Acid/Base Study Guide

Acid- pH <7 ; **Hydrogen** Donor ; Sour **taste**

Acids turn Blue litmus paper Red and is clear when mixed with phenolphthalein.

Examples-

Base pH >7 ; **Hydrogen** Acceptor ; Bitter **taste**

Bases turn Red litmus paper Blue and is pink when mixed with phenolphthalein.

Examples-

The acidity or basicity of a substance is measured on a pH scale. This scale ranges on a scale from 0-14. It is an acid if the pH < 7. It is a base if the pH > 7. It is neutral if the pH = 7.

The pH scale- pH is calculated using a logarithmic scale. If the pH is decreased by 1 unit, the substance is 10x more acidic. **pH = -log[H^+]**. The [H^+] can be found if you know the molarity of the acid. For bases, the **pOH** can also be calculated (pOH = -log[OH^-]. pH + pOH = 14.

Example- Identify the following substances as acidic or basic.

a. seawater (pH = 8.6) Basic
b. drain cleaner (pOH = 1) Basic
c. milk (pH = 6.3) Acidic
d. pure water (pOH = 7) Neutral

Reactions involving acids & bases

There are 2 common reactions involving acids & bases.

The first one is a reaction between an acid and a metal.

General Reaction: Acid + metal → Hydrogen gas + salt

This reaction is basically a single replacement reaction.

Example: Hydrochloric acid + zinc → Hydrogen gas + zinc chloride

$2HCl\ (aq) + Zn\ (s) \rightarrow H_2(g) + ZnCl_2$

The second reaction is called a **neutralization reaction.**

General reaction: Acid + Base → Salt + water

This reaction is basically a double replacement reaction

Example: Hydrochloric acid + sodium hydroxide → Sodium Chloride + water

$HCl(aq) + NaOH(aq) \rightarrow NaCl + H_2O(l)$

In a neutralization reaction, the acid and the base "cancel each other out" and the resulting product is neutral. This happens when **equal moles (not necessarily equal molarities or volumes)** of acids and bases are mixed. It is then "neutralized" and has a pH of 7.

Moles acid = moles base → (Molarity acid x volume acid) = (Molarity base x volume base)

$M_aV_a = M_bV_b$

Example: 90 mL of KOH is neutralized with 15 mL of 0.25 M HNO_3. What is the strength of the KOH? What is the pH of both solutions?

$M_aV_a = M_bV_b$ → (0.25 M x 15 mL) = (M_b x 90 mL) M_b = (0.25 x 15)/90 = 0.0417 M

pH of acid = -log(0.25) = 0.602

pH of base = 14-pOH of base = 14-1.38 = 12.6

pOH = -log(0.0417) = 1.38

pH AND pOH

Name _____

The pH of a solution indicates how acidic or basic that solution is.

pH range of 0 - 7 acidic
7 neutral
7-14 basic

Since $[H^+]\,[OH^-] = 10^{-14}$ at 25° C, if $[H^+]$ is known, the $[OH^-]$ can be calculated and vice versa.

$pH = -\log [H^+]$ So if $[H^+] = 10^{-6}$ M, pH = 6.
$pOH = -\log [OH^-]$ So if $[OH^-] = 10^{-8}$ M, pOH = 8.
Together, pH + pOH = 14.

Complete the following chart.

	$[H^+]$	pH	$[OH^-]$	pOH	Acidic or Basic
1.	10^{-5} M	5	10^{-9} M	9	Acidic
2.	10^{-7} M	7	10^{-7} M	7	Neutral
3.	10^{-10}	10	10^{-4} M	4	Basic
4.	10^{-2} M	2	10^{-12}	12	Acidic
5.	10^{-3}	3	10^{-11}	11	Acidic
6.	10^{-12}	12	10^{-2}	2	Basic
7.	10^{-9}	9	10^{-5} M	5	Basic
8.	10^{-11} M	11	10^{-3}	3	Basic
9.	10^{-1}	1	10^{-13}	13	Acidic
10.	10^{-6}	6	10^{-8}	8	Acidic

(3)

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pH AND pOH CONTINUED

Name _____

Calculate the pH of the solutions below.

1. 0.01 M HCl
   1×10^{-2}
   $pH = -\log(.01) = 2$

2. 0.0010 M NaOH
   1×10^{-3}
   $pOH = -\log(.001) = 3$
   $pH = 14-3 = 11$

3. 0.050 M $Ca(OH)_2$
   $.05\ M \times 2 = .1$
   1×10^{-1}
   $pOH = -\log(.1) = 1$
   $pH = 14-1 = 13$

4. 0.030 M HBr
   $pH = -\log(.03) = 1.52$

5. 0.150 M KOH
   $pOH = -\log(.150) = .823$
   $pH = 14 - .823 = 13.176$

6. 2.0 M $HC_2H_3O_2$ (Assume 5.0% dissociation.)
   $2.0M \times .05 = .1$
   $pH = -\log(.1) = 1$

7. 3.0 M HF (Assume 10.0% dissociation.)
   $3.0 \times .1 = .3$
   $pH = -\log(.3) = .552$

8. 0.50 M HNO_3
   $pH = -\log(.5) = .301$

9. 2.50 M NH_4OH (Assume 5.00% dissociation.)
   $2.5 \times .05 = .125$
   $pOH = -\log(.125) = .903$
   $pH = 14 - .903 = 13.097$

10. 5.0 M HNO_2 (Assume 1.0% dissociation.)
    $5.0M \times .1 = .5M$
    $pH = -\log(.5) = .301$

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Titrations Practice Worksheet

Find the requested quantities in the following problems:

1) If it takes 54 mL of 0.1 M NaOH to neutralize 125 mL of an HCl solution, what is the concentration of the HCl?

$M_A V_A = M_B V_B$

$(M_A)(125\,mL) = (.1\,M)(54)$

$M_A = .0432\ M$

2) If it takes 25 mL of 0.05 M HCl to neutralize 345 mL of NaOH solution, what is the concentration of the NaOH solution?

$M_A V_A = M_B V_B$

$(.05M)(25mL) = M_B(345mL)$

$M_B = .00362\ M$

3) If it takes 50 mL of 0.5 M KOH solution to completely neutralize 125 mL of hydrocyanic acid solution (HCN), what is the concentration of the HCN solution?

$M_A V_A = M_B V_B$

$(M_A)(125mL) = (.5M)(50mL)$

$M_A = .2M$

4) Can I titrate a solution of unknown concentration with another solution of unknown concentration and still get a meaningful answer? Explain your answer in a few sentences.

No, You need to know the concentration of at least one of the species.

Oxidation-Reduction Study Guide

An oxidation/reduction (RedOx) reaction involves the transfer of electrons. One species loses electrons while another species gains electrons. Take this example, a copper solution mixed with zinc metal will produce copper metal and a zinc solution.

$$Cu^{+2}(aq) + Zn(s) \rightarrow Cu(s) + Zn^{+2}(aq)$$

Copper gained 2 electrons to form copper metal while zinc metal lost 2 electrons and formed zinc ions. When a substance is oxidized, it loses electrons. When a substance is reduced,, it gains electrons. There are 2 easy acronyms to help remember this- **OIL RIG** and **LEO (the lion goes) GER.**

Oxidation Is Loss Reduction Is Gain **Loss Electrons Oxidation Gain Electrons Reduction**

The way we tell whether something is oxidized or reduced is by looking at a substances **OXIDATION NUMBER**. If a substance's oxidation # goes up, it is losing electrons and being oxidized. If a substance's oxidation # goes down, it is gaining electrons and is being reduced.

When a redox reaction is balanced, the number of electrons lost = the number of electrons gained.

Rules for determining oxidation #'s.

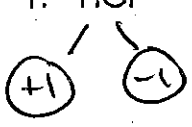
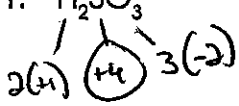
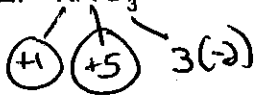
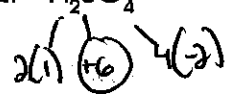
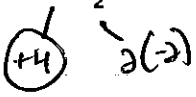
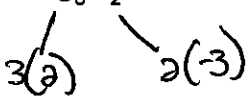
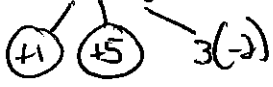
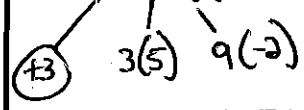
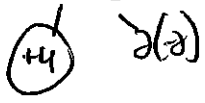
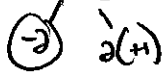
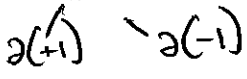
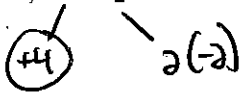
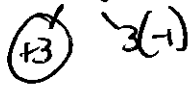
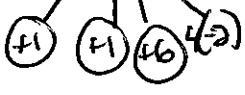
1. **All neutral elements have an oxidation # of zero.**
2. **Hydrogen (not H_2 gas) always has an oxidation # of +1.**
3. **Oxygen (not O_2) has an oxidation # of -2, except when in a peroxide, then it is -1.**
4. **Monatomic ions (Cl^-, Na^+, Ca^{+2}, Li^+) have an oxidation # equal to their charge.**
5. **In a neutral compound, the sum of all oxidation #'s in the compound has to add up to zero.**
   Example- $ZnCl_2$ **Zn has an oxidation # of +2 (Zn is a +2 charge)**
   Cl has an oxidation # of -1 (Cl is a -1 charge),
   But there are 2 Cl's, so you have to count the oxidation # twice. +2 + (-2) = 0
6. **In a polyatomic ion, the sum of all oxidation #'s for the ion has to add up to the charge of the ion.**
   Example- NO_3^{-1} **O has an oxidation # of -2, but there are 3 of them**
   N has to have an oxidation # of + 5
   5 + (-6) = -1
   The charge of the ion (-1) has to equal the oxidation #

   Example- NO_2^{-1} **O has an oxidation # of -2, but there are 2 of them**
   N has to have an oxidation # of + 3
   3 + (-4) = -1
   The charge of the ion (-1) has to equal the oxidation #

ASSIGNING OXIDATION NUMBERS

Name _____

Assign oxidation numbers to all of the elements in each of the compounds or ions below.

1. HCl 	11. H_2SO_3 
2. KNO_3 	12. H_2SO_4 
3. OH^- 	13. BaO_2 
4. Mg_3N_2 	14. $KMnO_4$ 
5. $KClO_3$ 	15. LiH 
6. $Al(NO_3)_3$ 	16. MnO_2 
7. S_8 	17. OF_2 
8. H_2O_2 	18. SO_3 
9. PbO_2 	19. NH_3 
10. $NaHSO_4$ 	20. Na 

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Name: Period: Date:

Oxidation Numbers and Oxidation & Reduction reactions- Section 19.1

1. Identify each of the following changes as either oxidation or reduction. Recall that e^- is the symbol for an electron.
   a. $I_2 + 2e^- \rightarrow 2I^-$ red
   b. $K \rightarrow K^+ + e^-$ ox
   c. $Fe^{2+} \rightarrow Fe^{3+} + e^-$ ox
   d. $Ag^+ + e^- \rightarrow Ag$ red

2. Identify what is oxidized and what is reduced in the following processes.
   a. $\overset{-1}{2Br^-} + \overset{0}{Cl_2} \rightarrow \overset{0}{Br_2} + \overset{-1}{2Cl^-}$ Br^- is oxidized Cl_2 is reduced
   b. $\overset{0}{2Ce} + \overset{+2}{3Cu^{2+}} \rightarrow \overset{0}{3Cu} + \overset{+3}{2Ce^{3+}}$ Ce is oxidized Cu^{+2} is reduced
   c. $\overset{0}{2Zn} + \overset{0}{O_2} \rightarrow \overset{+2\ -2}{2ZnO}$ Zn is oxidized O_2 is reduced
   d. $\overset{0}{2Na} + \overset{+1}{2H^+} \rightarrow \overset{+1}{2Na^+} + \overset{0}{H_2}$ Na is oxidized H^+ is reduced

3. Identify the oxidizing agent and the reducing agent in the following equation. Explain your answer.

$$\overset{0}{Fe(s)} + 2\overset{+1}{Ag^+}(aq) \rightarrow \overset{+2}{Fe^{2+}}(aq) + \overset{0}{2Ag(s)}$$

Fe(0 → +2) oxidized. Fe is red. Agent Ag^+ (+1 → 0) reduced Ag^{+1} is ox. agent.

4. Identify the oxidizing agent and the reducing agent in each reaction.
   a. $\overset{0}{Mg} + \overset{0}{I_2} \rightarrow \overset{+2\ -1}{MgI_2}$ Mg is red agent I_2 is ox agent.
   b. $\overset{+1\ -2}{H_2S} + \overset{0}{Cl_2} \rightarrow \overset{+1\ -1}{2HCl}$ Cl_2 is ox agent S^{-2} is red agent.

5. Determine the oxidation number of each element in the following compounds.
   a. $NaClO_4$ Na(+1) Cl(+7) O (4)(-2)
   b. $AlPO_4$ Al(+3) P(+5) O 4(-2)
   c. HNO_2 H(+1) N(+3) O 2(-2)
   d. NH_4^+ N(-3) H 4(+1)
   e. AsO_4^{3-} As(+5) O (4)(-2)
   f. CrO_4^{2-} Cr(+6) O (4)(-2)
   g. NH_3 N(+3) H(3)(-1)
   h. KCN K(+1) C(+2) N(-3)
   i. N_2H_4 N 2(-2) H 4(1)

6. Write the equation for the reaction of iron metal with hydrobromic acid to form iron (III) bromide and hydrogen gas. Determine the change in oxidation number for the element that is reduced and the element that is oxidized.

$$\overset{0}{Fe} + \overset{+1\ -1}{HBr} \longrightarrow \overset{+3\ \ -1}{FeBr_3} + \overset{0}{H_2}$$

Fe: 0 → +3 - lose $3e^-$ - oxidation

H: +1 → 0 Gain $1e^-$ - reduction

Br: -1 → -1