20. How many mL of 0.152 M HNO_3 would be needed to titrate 11.21 mL of 0.254 M KOH?

$NV = NV$ $(1)(0.152)(V) = (1)(0.254)(11.21)$ $V = 18.7\ mL$

21. A HCl solution has a pH = 2.55. What will be the molarity of the solution?

$[H^+] = 10^{-2.55} = 0.0028\ M$ $HCl \rightarrow H^+ + Cl^-$

22. How many ml of 0.152 M H_3PO_4 would be needed to titrate 11.21 mL of 0.254 $Mg(OH)_2$?

$NV = NV$ $(3)(0.152)(V) = (2)(0.254)(11.21)$ $V = 12.5\ mL$

23. Write the balanced equation for the reaction in #22. (H_2O)

$2\ H_3PO_4 + 3\ Mg(OH)_2 \rightarrow Mg_3(PO_4)_2 + 6\ H(OH)$

24. 15.0 g of nitrous acid (a weak acid) is dissolved into 750. mL of solution. What is the pH of this solution? (look up Ka on WS 10.9)

$15.0\ g \times \frac{1\ mol}{47\ g} = 0.319\ mol$

$M = \frac{mol}{L} = \frac{0.319\ mol}{0.750\ L} = 0.425\ M$

$HNO_2 \rightleftarrows H^+ + NO_2^-$

Start: 0.425 M ⇄ 0 M + 0 M

equilibrium: 0.425 − x ⇄ x + x

$K_a = \frac{[H^+][NO_2^-]}{[HNO_2]}$

$4.5 \times 10^{-4} = \frac{(x)(x)}{(0.425 - x)}$

$1.91 \times 10^{-4} - 4.5 \times 10^{-4} x = x^2$

$x^2 + 4.5 \times 10^{-4} x - 1.91 \times 10^{-4} = 0$

$x = 1.36 \times 10^{-2}\ M$

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

$= -\log [1.36 \times 10^{-2}]$

$= 1.9$

$H_2CO_{3(aq)} \rightleftharpoons CO_{2(g)} + H_2O_{(l)} + 13\ KJ$

25. List 5 ways to increase the amount of water produced by the above reaction.

- remove H_2O as it's produced.
- remove CO_2
- add H_2CO_3
- remove heat (cool down)
- decrease pressure

26. What is the oxidation # of **sulfur** in each of the following?

CaS_2O_3 +2 H_2SO_3 +4 SO_3 +6 SO_4^{2-} +6

Identify what is **oxidized**, what is **reduced**, the **oxidizing agent**, and the **reducing agent** for each:

27. $\overset{+5}{HNO_3} + \overset{+1}{Cu_2O} \rightarrow \overset{+2}{Cu(NO_3)_2} + \overset{+2}{NO} + H_2O$

(HNO_3: oxidizing agent; Cu_2O: reducing agent; Cu: oxidized; N: reduced)

28. $\overset{0}{I_2} + \overset{+5}{HNO_3} \rightarrow \overset{+5}{HIO_3} + \overset{+4}{NO_2}$

(I_2: reducing agent; HNO_3: oxidizing agent; I: oxidized; N: reduced)

29. Consider a voltaic cell containing Ag in a solution of AgCl and Pb in a solution of $PbCl_2$.

a. Write the half-reaction for the anode (oxidation): $Pb \rightarrow 2e^- + Pb^{+2}$

b. Write the half-reaction for the cathode (reduction): $Ag^+ + 1e^- \rightarrow Ag$

c. Use the shorthand method to represent this cell: $Pb \mid Pb^{+2} \parallel Ag^+ \mid Ag$

d. Calculate the cell potential: cathode − anode

$0.80 - (-0.13) = 0.93\ V$