

Normality Primer

Introduction

Normality, equivalents and equivalent weight are all related terms typically used in titrations when the titration reaction is unknown or just not used. Consequently, definitions for these terms vary depending on the type of chemical reaction that is being used for the titration. The two most common types of reactions for which normality is used are acid-base reactions and redox (reduction-oxidation) reactions.

The basic unit for normality related conventions (terms and calculations) is the *equivalent*. Equivalents (eq) are comparable to moles and used to stoichiometrically relate one substance to another. Normality (N) is comparable to molarity and has the units of equivalents per liter. Equivalent weight (eq wt) is comparable to molecular weight (molar mass) with units of grams per mole. When these terms are used, stoichiometric calculations can be performed without reference to or direct knowledge of the chemical reaction.

Modern academics avoid the use of these terms thus they are rarely found in modern textbooks. However, many industries continue to use them, therefore it is an advantage to know its use and terminology.

If you work completely with normality, equivalents and equivalent weights, you do not need to know the chemical reactions. However, if you need to convert from standard conventions (molarity, moles and molecular weight), you will.

Acid-Base Reactions

In an acid-base reaction, acids react with bases. If either the acid or base is strong, the reaction will go to completion and be can used in an acid-base titration. Under the Brønsted-Lowry theory, acids are defined as proton (H^+) donors and bases as proton acceptors. Thus an acid-base reaction is a proton transfer. When using normality conventions, an equivalent is defined as the moles of protons (H^+) accepted or donated by the substance. These are moles that react in a complete acid-base reaction (like a titration reaction) rather than the hydrolysis reactions associated with equilibrium. Therefore, whether the acid or base is strong or weak is not relevant to determination of equivalents.

Equivalent (eq) = moles of H^+ accepted or donated.

Examples: 1.0 mol HCl = 1.0 eq HCl
1.0 mol H_3PO_4 = 3.0 eq H_3PO_4
1.0 mol KOH = 1.0 eq KOH
1.0 mol $Ba(OH)_2$ = 2.0 eq $Ba(OH)_2$
1.0 mol Na_2CO_3 = 2.0 eq Na_2CO_3

The other relationship definitions follow from the definition of equivalent where normality (N) is equivalents/liter = eq/L and equivalent weight (eq wt) = grams/equivalent = g/eq.

Normality (N) = equivalents/liter = eq/L

Examples: 1.0 M HCl = 1.0 N HCl
1.0 M H_2S = 2.0 N H_2S
1.0 M KOH = 1.0 N KOH
1.0 M $Fe(OH)_3$ = 3.0 N $Fe(OH)_3$
1.0 M Na_2CO_3 = 2.0 N Na_2CO_3

Equivalent weight (eq wt) = grams/equivalent = g/eq

Examples: 204.23 g/mol KHP = 204.23 g/eq KHP
105.99 g/mol Na₂CO₃ = 53.0 g/eq Na₂CO₃
106.88 g/mol Fe(OH)₃ = 35.63 g/eq Fe(OH)₃

Because the acid donates protons to the base, in an acid-base reaction, equivalents of acid are always equal to equivalents of base. This relationship is used instead of the balanced chemical equation.

equivalents of acid = equivalents of base (eq acid = eq base)

Putting together the relationships between the equivalents of acid and base and the definition of normality, other useful relationships can be obtained. One of the most convenient is the relationship between normality and volume in a titration. Where the normality of the acid (N_a) times the volume of the acid (V_a) is equal to the normality of the base (N_b) times the volume of the base (V_b).

This is often confused with the dilution equation and care must be taken to use *normality*, not molarity, when applying it to titrations. However, like the dilution equation, any volume unit can be used provided it is the same for both the acid and the base.

$$N_a V_a = N_b V_b$$

Example: How many milliliters of 0.123 N NaOH is required to titrate 25.00 mL of 0.321 N H₃PO₄?

$$(0.321 \text{ N})(25.00 \text{ mL}) = (0.123 \text{ N})(V_b)$$

$$65.2 \text{ mL} = V_b$$

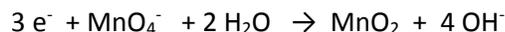
Reduction-Oxidation (Redox) Reactions

Reduction-oxidation reactions are reactions where electrons are transferred. Electrons are gained in the *reduction* half and lost in the *oxidation* half. The reactant that is oxidized is the *reducing agent* (since oxidation causes reduction) and the reactant that is reduced is the *oxidizing agent*. Consequently, the oxidizing and reducing agents are the two major reactants. In titrations, one is normally the analyte and the other the standard. However, redox reactions are normally done in acidic or basic conditions, so an excess of acid or base is generally required, but, as long as it is in excess, it does not directly affect the calculations. Since redox reactions are essentially the transfer of electrons, equivalents are defined as moles of electrons lost or gained.

To know how many electrons are gained or lost, redox reactions require knowledge of the half reaction associated with the reducing or oxidizing agent. These reactions are not easily predictable and generally chemists look them up in tables of reduction potentials. However, one must be particularly careful to choose the correct half-reaction as it can vary with conditions. For example, under acidic conditions, permanganate reacts as:



But under basic conditions, permanganate reacts as:



Consequently, the way equivalents are measured will depend not only on the substance's identity, such as permanganate (the oxidizing agent), but also on experimental conditions.

equivalents = moles of electrons gained or lost

Examples:



As with acid-base reactions, the other definitions follow from the definition of equivalents and are the same based on equivalents where normality (N) = equivalents/liter = eq/L and equivalent weight (eq wt) = grams/equivalents = g/eq.

Redox reactions are normally written as net ionic equations, which should not affect normality and equivalent calculations. However, equivalent weights will require knowledge of the entire compound...since you can't weigh an ion without its counterion.

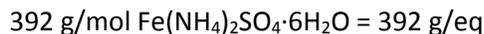
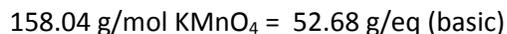
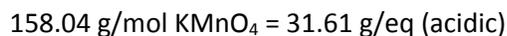
Normality (N) = equivalents/liter = eq/L

Examples:



Equivalent weight (eq wt) = grams/equivalents = g/eq

Examples:



In a redox reaction, the moles of electrons lost in the oxidation reaction by the reducing agent are gained in the reduction reaction by the oxidizing agent.

equivalents of reducing agent = equivalents of oxidizing agent

As with acid-base reactions, this relationship is used in place of the balanced chemical reaction for stoichiometric calculations. You can also use volume to relate the normality of the reducing agent (N_r) to normality of the oxidizing agent (N_o) in the same way normality of acid and base are related.

$$N_r V_r = N_o V_o$$

Example: How many milliliters of 0.345 N KMnO_4 are required to titrate 25.00 mL of 0.876 N Cu^{2+} in 50.00 mL of 1.5 N HCl?

$$(25.00 \text{ mL})(0.876 \text{ N}) = (0.345 \text{ N})(V)$$

$$63.5 \text{ mL} = V$$

Hint: The only sneaky part to this problem is knowing that the permanganate reacts with the copper rather than the acid. Acid is a necessary condition for the reaction, but added in excess so it does not affect the calculations. This is pretty typical of redox titrations.

Other examples

Equivalents are also sometimes used for water analysis where they are defined as the amount of charge on an ion. Thus, 1.0 mol of Na^+ is 1.0 eq and 1.0 mol Ca^{2+} = 2.0 eq.

Equivalents can be used to relate reactants in any type of reaction, but the definition will be specific to that type of reaction. Consequently, that definition must be known to do the calculations effectively.

Problems

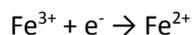
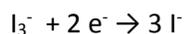
Acid- Base Titrations

- Convert the following to equivalents.
 - 0.12 mol H_2SO_4
 - 0.34 mol NaHCO_3
 - 0.56 mol $\text{Cr}(\text{OH})_3$
- Convert the following to moles.
 - 0.98 eq H_2S
 - 0.763 eq CH_3COOH
 - 0.5647 eq $\text{Ba}(\text{OH})_2$
- Convert the following to normality.
 - 0.54 M NH_3
 - 0.43 M $\text{H}_2\text{C}_2\text{O}_4$
 - 0.32 M H_3AsO_4
- Convert the following to molarity.
 - 0.211 N CH_3NH_2
 - 0.1951 N $\text{Ca}(\text{OH})_2$
 - 0.334 N $\text{C}_6\text{H}_5\text{COOH}$
- What is the equivalent weight of each of the following substances?
 - NH_3
 - $\text{Ca}(\text{OH})_2$
 - H_3AsO_4
 - H_2S
- How many milliliters of 0.341 N H_2SO_4 are required to react with 30.0 mL of 0.333 N $\text{Cr}(\text{OH})_3$?
- What is the normality of a solution of unknown base if 23.45 mL of 0.1236 N HCl are required to neutralize 33.50 mL of it?
- To standardize a solution of HCl, 0.1423 g of Na_2CO_3 (eq wt = 53.00) was dissolved in 50 mL of water and titrated with 27.82 mL of HCl. What is the normality of the $\text{HCl}(aq)$?
- To standardize a solution of NaOH, 0.5531 g of KHP (eq wt = 204.23) was dissolved in 50 mL of water. This solution required 31.11 mL of $\text{NaOH}(aq)$ for neutralization. What is the normality of the $\text{NaOH}(aq)$?

10. A 0.9932 g sample of limestone was titrated with 15.67 mL of 0.113 N HCl, what is the percent of calcium carbonate in the sample?
11. 27.44 mL of 0.222 N Ba(OH)₂ was required to neutralize all the benzoic acid (C₆H₅COOH) in a 1.224 g sample of organic material. What was the percent benzoic acid in the sample?
12. The citric acid in a 0.541 g vitamin tablet was dissolved in 20.00 mL of 1.021 N NaOH. The excess base was titrated with 9.21 mL of 0.223 N HCl. How many mg of citric acid (eq wt = 64.0) was in the vitamin tablet?
13. A 1.113 g sample of antacid was dissolved in 50.00 mL of 0.555 N HCl. The excess HCl was titrated with 10.2 mL of 0.033 N NaOH. What is the power of the antacid in meq/g?
14. A 0.1298 g sample of pure acid was titrated with 42.11 mL of 0.1234 N NaOH. What is the equivalent weight of the acid?
15. A 0.3349 g sample of pure base was titrated with 33.33 mL of 0.1673 N H₂SO₄. What is the equivalent weight of the base?

Redox Titrations

Refer to the following half-reactions, as needed, to answer the questions below about redox reactions.



- Convert the following to equivalents.
 - 0.222 mol IO₃⁻
 - 0.045 mol Fe²⁺
 - 0.31 mol I⁻
- Convert the following to moles.
 - 0.55 eq MnO₄⁻ (acidic conditions)
 - 0.0781 eq Zn
 - 0.211 eq IO₃⁻

3. Convert the following to normality.
 - a. 0.101 M Fe^{2+}
 - b. 0.0056 M MnO_4^- (basic conditions)
 - c. 0.207 M I_3^-

4. Convert the following to molarity.
 - a. 0.440 N Mn^{2+}
 - b. 1.03 N IO_3^-
 - c. 0.224 N NO_3^-

5. How many milliliters of 0.134 N permanganate ion is needed to titrate 25.00 mL of 0.223 N iron(II) ion?

6. What is the normality of an iodate solution if 35.44 mL of 0.0111 N thiosulfate is required to titrate 30.00 mL of the iodate solution?

7. How many grams of zinc metal are needed to reduce 25.00 mL of 0.123 N nitric acid?

8. What is the percent of iron in a 1.2299 g sample that requires 14.35 mL of 0.075 N acidic permanganate to titrate it from iron (II) to iron(III)?

9. What is the normality of a basic solution of permanganate ion if 15.00 mL of the solution requires 16.43 mL of 0.143 N oxalate ion for a complete reaction?

10. What is the normality of a solution of copper(II) ion that requires 21.33 mL of 0.236 N silver nitrate to titrate 25.00 mL of the copper solution?