

PERIODIC TABLE WITH CHEMISTRY FORMULAS

CHEMISTRY FORMULAS

ATOMS AND SUBATOMIC PARTICLES

A. Atomic number: The number of protons in an atom of a given element. In the periodic table, elements are ordered by ascending atomic number.

Ex: $^{12}_{\text{C}}$ = 6 = atomic number

B. Mass number: The number of protons plus the number of neutrons in an atom of an element, measured in atomic mass units (amu). 1 amu = mass of one proton or neutron = 1.66×10^{-27} kg.

Ex: $^{12}_{\text{C}}$ = 12 = mass number

Name	Symbol	Mass	Charge
Proton	p^+	1 amu	+1
Neutron	n^0	1 amu	0
Electron	e^-	1/1837 amu	-1

C. Isotopes: Variant atoms of an element that possess different numbers of neutrons, and therefore different mass numbers.

Ex: Boron-10: $^{10}_{\text{B}}$ Boron-11: $^{11}_{\text{B}}$

D. Calculating atomic mass: An element's atomic mass number on the periodic table is a weighted average of that element's isotopes.

Ex:

Isotope	Abundance	Mass #	Isotopic Mass
Boron-10	19.78%	10	10.013 amu
Boron-11	80.22%	11	11.009 amu

Atomic mass = $\frac{(19.78\%)(10.013 \text{ amu}) + (80.22\%)(11.009 \text{ amu})}{100} = 10.81 \text{ amu}$

MOLECULAR MASS

A. Avogadro's number: 6.022×10^{23} 6.022 $\times 10^{23}$ atoms or molecules = 1 mole of that element or substance

B. The atomic mass number of an element equals the mass (in grams) of one mole of that element.

Ex: One mole of boron has a mass of 10.81 grams.

MOLECULAR FORMULA AND EMPIRICAL FORMULA

A. Molecular formula: The chemical formula of a compound that gives the actual number of atoms of each element in a molecule of the compound.

Ex: The molecular formula for glucose is $\text{C}_6\text{H}_{12}\text{O}_6$.

B. Empirical formula: A "lowest terms" expression of a chemical formula.

Ex: The empirical formula for glucose is CH_2O .

BALANCING CHEMICAL EQUATIONS

- Check for diatomic molecules such as H_2 and O_2 .
- Balance metals.
- Balance nonmetals.
- Balance oxygen.
- Balance hydrogen.
- You may change only coefficients, never the formula subscripts! Recount all atoms to be sure you have equal numbers on both sides of the equation.
- Rewrite the equation so that coefficients are the smallest whole numbers possible.

PERCENTAGE AND PPM

A. Percentage composition: Tells how much mass an object or element contributes to a compound:

$$\text{Percentage composition} = \frac{\text{mass of element}}{\text{mass of molecule}} \times 100$$

B. Parts per million (ppm): In a compound or solution, the number of grams of a

particular component that would be present in a one million gram sample of the compound or solution.

TEMPERATURE CONVERSION

A. Converting between Fahrenheit and Celsius

$$T_{\text{Fahrenheit}} = \frac{9}{5}C(T_{\text{Celsius}}) + 32^\circ\text{F}$$

$$T_{\text{Celsius}} = \frac{5}{9}(T_{\text{Fahrenheit}} - 32^\circ\text{F})$$

B. Kelvin: The temperature 0 K, **absolute zero**, is the theoretical lowest possible temperature, at which all molecular activity ceases.

$$0 \text{ K} = -273.15^\circ\text{C} = -459.67^\circ\text{F}$$

$$K = ^\circ\text{C} + 273.15 \quad ^\circ\text{C} = K - 273.15$$

IDEAL GAS LAWS

P = pressure

V = volume (expressed in liters)

T = temperature (expressed in Kelvins)

Standard temperature and pressure (STP):

$$T = 0^\circ\text{C} = 273 \text{ K}$$

$$P = 1 \text{ atm} = 760 \text{ mmHg} = 101.3 \text{ kPa}$$

A. Boyle's law: $PV = \text{constant}$

$$P_1V_1 = P_2V_2$$

B. Charles' law: $\frac{V}{T} = \text{constant}$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

C. Gay-Lussac's law: $\frac{P}{T} = \text{constant}$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

D. Combined gas law: $\frac{PV}{T} = \text{constant}$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

E. Ideal gas law: $PV = nRT$

n = number of moles; R = ideal gas constant

$$R = 8.31 \frac{\text{L} \cdot \text{kPa}}{\text{mol} \cdot \text{K}} = 62.4 \frac{\text{L} \cdot \text{mmHg}}{\text{mol} \cdot \text{K}} = 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$$

F. Dalton's law of partial pressure:

$$\frac{n_{\text{gas } 1}}{n_{\text{total}}} = \frac{P_{\text{gas } 1}}{P_{\text{total}}}$$

$$\frac{n_{\text{gas } 2}}{n_{\text{total}}} = \frac{P_{\text{gas } 2}}{P_{\text{total}}}$$

$$P_{\text{total}} = P_{\text{gas } 1} + P_{\text{gas } 2}$$

SOLUBILITY

A. Molarity (M): Measures the number of moles (n) of solute per liters of solution:

$$M = \frac{n_{\text{solute}}}{\text{liters of solution}}$$

B. Molality (m): Measures the number of moles (n) of solute dissolved in 1 kg of solvent:

$$m = \frac{n_{\text{solute}}}{\text{kg solvent}}$$

C. Dilution: The relationship between initial volume (V_i) and final volume (V_f) is inversely proportional to the relationship between initial molarity (M_i) and final molarity (M_f):

$$\frac{V_f}{V_i} = \frac{M_i}{M_f} \quad V_iM_i = V_fM_f$$

COLLIGATIVE PROPERTIES

A. Raoult's law (vapor-pressure lowering):

When solute B is added to pure solvent A,

$$P_{\text{actual}} = X_A P_A^0$$

where P_{actual} is the vapor pressure of A after B is added, X_A is the mole fraction of A, and P_A^0 is the vapor pressure of A above pure solvent A.

B. Freezing-point depression:

$$\Delta T_f = K_f m,$$

where ΔT_f is the amount the freezing point is lowered, K_f is a proportionality constant specific to the solvent, and m is the molality of the solution.

C. Boiling-point elevation:

$$\Delta T_b = K_b m,$$

ΔT_b is the amount the boiling point is elevated, K_b is a proportionality constant specific to the solvent, and m is the molality of the solution.

D. Osmotic pressure:

$$\pi = MRT,$$

where π is the osmotic pressure, M is the molarity of the solution, R is the ideal gas constant, and T is the temperature in Kelvins.

CHEMICAL EQUILIBRIA

A. Equilibrium constant expression

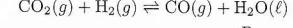
Method using molar concentration:

For the reaction $aA + bB \rightleftharpoons cC + dD$:

$$\text{Equilibrium constant: } K_{\text{eq}} = \frac{[C]^c[D]^d}{[A]^a[B]^b}.$$

Method using gas partial pressures:

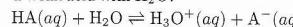
For the reaction



$$\text{Equilibrium constant: } K_{\text{eq}} = \frac{P_{\text{CO}}}{P_{\text{CO}_2}P_{\text{H}_2}}.$$

B. Acid-Base equilibrium

1. Acid dissociation reactions: The reaction of a weak acid with H_2O :

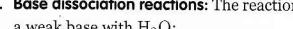


Acid dissociation constant (equilibrium constant for the reaction):

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

The greater the K_a value, the stronger the acid.

2. Base dissociation reactions: The reaction of a weak base with H_2O :



Base dissociation constant (equilibrium constant for the reaction):

$$K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}$$

The greater the K_b value, the stronger the base.

3. Dissociation of water: The self-ionization of water:



$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

4. Conjugate acid-base pair:

$$K_aK_b = K_w = 1.0 \times 10^{-14}$$

5. pH and pOH: pH measures $[\text{H}^+]$; pOH measures $[\text{OH}^-]$.

— **Acids:** $\text{pH} < 7$ and $\text{pOH} > 7$

— **Bases:** $\text{pH} > 7$ and $\text{pOH} < 7$

— **Water:** $\text{pH} = \text{pOH} = 7$

Calculating pH: $\text{pH} = -\log_{10}[\text{H}_3\text{O}^+]$

Calculating pOH: $\text{pOH} = -\log_{10}[\text{OH}^-]$

For all aqueous solutions:

$$[\text{OH}^-][\text{H}_3\text{O}^+] = 1 \times 10^{-14}$$

$$\text{pH} + \text{pOH} = 14$$

REACTION RATES

For the reaction $aA + bB \rightarrow cC + dD$:

Rate = $k[A]^m[B]^n$, where k = rate constant;

m = rate order with respect to A; n = rate order with respect to B.

Ex: If the rate law is $k[A]^2[B]$, the reaction is second order with respect to A and first order with respect to B.

Ex: If the rate law is $k[A]^0 = k$, the reaction is

zero order with respect to A, meaning the reaction rate is independent of [A].

THEORETICAL AND PERCENTAGE YIELD

Percentage yield = $100 \times \frac{\text{Actual yield}}{\text{Theoretical yield}}$

THERMODYNAMICS

A. Specific heat:

$$q = mc\Delta T$$

where q is the heat absorbed by the object, m is the mass of the object, ΔT is the change in temperature, and c is the specific heat of the substance.

B. Enthalpy (H):

$$\Delta H = H_{\text{product}} - H_{\text{reactants}}$$

$\Delta H < 0$ for exothermic reactions

$\Delta H > 0$ for endothermic reactions

C. Entropy (S):

$$\Delta S = S_{\text{product}} - S_{\text{reactants}}$$

D. Gibbs free energy (G):

$$\Delta G = \Delta H - T\Delta S$$

$\Delta G = 0$ when at equilibrium

$\Delta G < 0$ when forward process is spontaneous

$\Delta G > 0$ when reverse process is spontaneous

REDOX REACTIONS AND ELECTROCHEMISTRY

A. Standard reduction potential of a cell:

$$E = E_{\text{cell}}^{\circ} = E_{\text{reduction}} + E_{\text{oxidation}}$$

B. Nernst equation:

$$E = E_{\text{cell}}^{\circ} - \frac{RT}{nF} \ln Q = \frac{0.0592 \text{ V}}{n} \log_{10} K_{\text{eq}}$$

where R is the universal gas law constant, T is in Kelvins, n is the number of electrons in the half-reaction, F (Faraday's constant) = 96485 C mol⁻¹, and Q is the reaction quotient.

C. Standard cell voltage at equilibrium:

$$E_{\text{cell}}^{\circ} = \frac{RT}{nF} \ln K_{\text{eq}} = \frac{0.0592 \text{ V}}{n} \log_{10} K_{\text{eq}}$$

where n is the total number of electrons transferred in the redox reaction.

D. Thermodynamics:

$$\Delta G^{\circ} = -nFE_{\text{cell}}^{\circ}$$

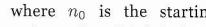
$$-nFE_{\text{cell}}^{\circ} = -RT \ln K_{\text{eq}}$$

RADIOACTIVE DECAY

A. Particle types

Alpha: ${}^{4}\text{He}$ Beta: ${}^{-1}\text{e}$ Gamma: ${}^{0}\gamma$ Positron: ${}^{0}\text{e}$ Neutron: ${}^{1}\text{n}$

Ex: Alpha decay:



B. Exponential decay:

$$n = n_0 e^{-\lambda t}$$

where n_0 is the starting number of undecayed nuclei, λ is the decay constant, and t is the time.

SIGNIFICANT FIGURES

A. Types of significant figures

— Any nonzero digits in a number

— Zeros appearing between two nonzero digits in a number

— Zeros appearing at the end of a number, but only if that number contains a decimal point

B. Multiplication/division: The result of the calculation should have the same number of significant figures as the least precise number being multiplied or divided.

C. Addition/subtraction: The result of the calculation should have the same number of decimal places as the least precise number being added or subtracted.