Thermochemical Equations for Phase Changes

Thermochemical equations can also be written for __physical__ changes!

$$H_2O(s) \rightarrow H_2O(l) \qquad \Delta H = +6.02 \text{ kJ/mol}$$

$$H_2O(l) \rightarrow H_2O(s) \qquad \Delta H = -6.02 \text{ kJ/mol}$$

Physical changes (__phase__ changes, or changes of __state__) are either endo- or exothermic:

- Endothermic physical changes

    1. Require an addition of heat/energy to overcome __IMFs__ (intermolecular forces) and change state (just like ending a bond, "end"ing an IMF is always "end"othermic!)
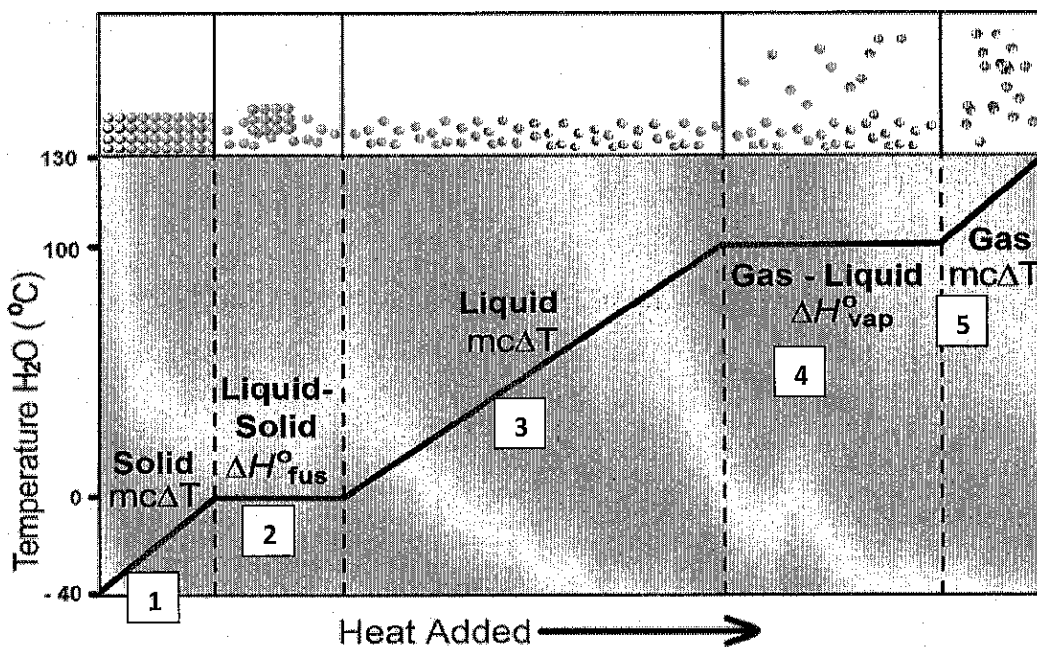
        - Solid → Liquid (Melting)
        - Liquid → Gas   (Vaporization)
        - Solid → Gas   (Sublimation)

- Exothermic physical changes

    1. Heat/energy is released as IMFs form

        - Gas → Liquid   (Condensing)
        - Liquid → Solid (Freezing)
        - Gas → Solid   (Deposition)

Heating Curve of Water

When heating a solid, liquid, or gas (1, 3, and 5):

- KE / Temp increases
- PE is constant
- Constant state

When changing phase in the forward direction (2 and 4):

- KE / Temp is constant
- PE increases
- Multiple states

Energy Changes

1. __Inclined__ (slanted) **sections** (1, 3, and 5): Temperature (thus __KE__) changing

    → Use mCΔT

2. __Flat__ (plateau) **sections** (2 and 4): 3 Ps: **P**lateau, **P**hase change, and **P**otential Energy Change

    → __NO__ change in temperature = KE constant

    → Use $q = n \Delta H$

How to Calculate Energy Needed to Change Temperature or Phase

1. **Temperature changes:** $\quad q = m\,C\,\Delta T$

    This **starts with grams** and gives **heat in Joules**

    Note: is your specific heat a *molar* heat capacity? (C_m = molar heat capacity = **J/mol** K)

    → $\quad$ Use $q = n\,C_m\,\Delta T$

2. **Phase changes:** TWO Options!

    - Melting $\qquad q = n\,\Delta H_{fus}$
    - Freezing $\qquad q = -n\,\Delta H_{fus}$

        n = moles of substance

        ΔH_{fus} = enthalpy of fusion

    - Vaporizing $\qquad q = n\,\Delta H_{vap}$
    - Condensing $\quad q = -n\,\Delta H_{vap}$

        n = moles of substance

        ΔH_{vap} = enthalpy of vaporization

    These **start with moles** and give **heat in kilojoules**

3. Once you know the heat at each step of the process, add them together to find the heat of the process, but **make sure you are adding the same units of energy!** (usually kJ +kJ)

Examples: Determine the amount of heat gained (+) or lost (−) during each of the following changes.

1. Melting 55.8 g of Ti at 1677°C , ΔH_{fus} (Ti) = 18.8 KJ/mol

$$q = n\,\Delta H$$

$$= \left(55.8\ \cancel{g} \times \frac{1\ mol\ Ti}{47.87\ g}\right)\left(18.8\ \frac{KJ}{mol}\right) = \boxed{21.9\ KJ}$$

2. Converting 45.0 g of water at 20.0°C to steam at 115°C $\quad C_{H_2O_{(\ell)}} = 4.18\ \frac{J}{g°C}$, $C_{H_2O_{(g)}} = 2.02\ \frac{J}{g°C}$

① heat $H_2O_{(\ell)}$: 20.0°C → 100°C $\Rightarrow q = mC\Delta T$

$$q = (45.0\ g)(4.18\ \tfrac{J}{g°C})(100.0°C - 20.0°C) = 15{,}048\ J = 15.048\ KJ$$

② boil $H_2O_{(\ell)} \to H_2O_{(g)}$: $\Delta H_{vap}(H_2O) = 40.7\ kJ/mol$

$$q = n\,\Delta H = \left(45.0\ g \times \frac{1\ mol\ H_2O}{18.016\ g}\right)\left(40.7\ \tfrac{KJ}{mol}\right) = 101.66\ KJ$$

③ heat $H_2O_{(g)}$: 100 − 115°C $\Rightarrow q = mC\Delta T$

$$q = (45.0\ g)(2.02\ \tfrac{J}{g°C})(115 - 100.) = 1363.5\ J = 1.3635\ KJ$$

ADD: $15.048 + 101.66 + 1.3635 = \boxed{118\ KJ}$
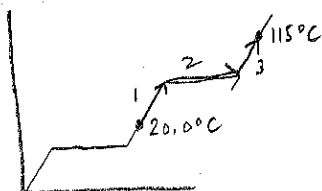
3. Calculate the energy absorbed when heating 6.9 g of solid aluminum from 32°C to 320°C. The melting point of aluminum is 660°C, ΔH_{fus} = +10.8 kJ/mol, and the specific heat capacity of Al(s) is 0.903 J/mol °C.

$$q = \left(6.9g \times \frac{1 mol}{26.98 g}\right)\left(0.903 \frac{J}{mol \cdot C}\right)(320. - 32) = \boxed{67 J}$$

$$q = n C \Delta T !$$

4. Calculate the energy absorbed when melting 27.3 g of Al at 660°C. The melting point of aluminum is 660°C, ΔH_{fus} = 10.8 kJ/mol, and the specific heat capacity of Al(s) is 0.903 J/mol °C.

$$q = n \Delta H = \left(27.3g \times \frac{1 mol\ Al}{26.98\ g}\right)\left(10.8 \frac{kJ}{mol}\right) = \boxed{10.9\ KJ}$$

5. When water vaporizes at its normal boiling point, its ΔH_{vap} = +40.7 kJ/mol. A sample of water vapor was condensed at 25°C and 1.05 atm. Calculate the volume of water that was condensed, in liters, if 620. kJ of energy was released during the process.

$$q = n \Delta H$$

$$-620. KJ = -n_{H_2O}\left(40.7 \frac{kJ}{mol}\right) \Rightarrow n = \frac{-620. KJ}{-40.7 kJ/mol} = 15.2\ mol$$

$$PV = nRT \Rightarrow V = \frac{nRT}{P} = \frac{(15.2\ mol)(0.08206 \frac{L \cdot atm}{mol \cdot K})(25 + 273)}{1.05\ atm} = \boxed{355\ L}$$

Part II: Determine the amount of heat gained (+) or lost (-) during each of the following changes. (For water, ΔH_{fus} = +6.02 kJ/mol, ΔH_{vap} = +40.7 kJ/mol, C_{solid} = 2.06 J/g °C, C_{liquid} = 4.18 J/g °C, and C_{gas} = 2.02 J/g °C).

6. 220.0 g of ice at −35.0°C is converted to liquid water at 50.0°C.

1.) heat $H_2O_{(s)}$: −35.0°C to 0.0°C $\Rightarrow q = m C \Delta T$
$$q = (220.0g)(2.06 \frac{J}{g \cdot c})(0.0 - (-35.0)) = 15,862\ J = 15.862\ kJ$$

2) melt $H_2O_{(s)} \rightarrow H_2O_{(l)}$ $\Rightarrow q = n \Delta H_{fus}$
$$q = \left(220.0g \times \frac{1 mol\ H_2O}{18.016g}\right)\left(6.02 \frac{kJ}{mol}\right) = 73.512\ KJ$$

3) heat $H_2O_{(l)}$: 0.0°C to 50.0°C $\Rightarrow q = m C \Delta T$
$$q = (220.0g)(4.18\ J/g°c)(50.0 - 0.0) = 45,980 = 45.98\ kJ$$

Add: 15.862 kJ + 73.512 kJ + 45.98 kJ = $\boxed{135.4\ kJ}$

7. 5.00 g of steam at 155°C is converted to liquid water at 100.0°C

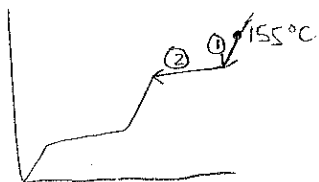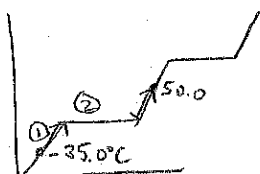1) cool $H_2O_{(g)}$: 155°C $\rightarrow$ 100.0°C $\Rightarrow q = m C \Delta T$
$$q = (5.00 g)(2.02 \frac{J}{g°c})(100. - 155) = -555.5\ J = -0.5555\ kJ$$

2) condense $H_2O_{(g)} \rightarrow H_2O_{(l)}$ $\Rightarrow q = -n \Delta H_{vap}$
$$q = -\left(5.00g \times \frac{1 mol\ H_2O}{18.016g}\right)\left(40.7 \frac{kJ}{mol}\right) = -11.296\ kJ$$

Add: −0.5555 kJ + (−11.296 kJ) = $\boxed{-11.9\ kJ}$

Multiple Choice Practice FTW!

8. Aluminum metal can be recycled from scrap metal by melting the metal to evaporate impurities.

Temperature (°C)

Time
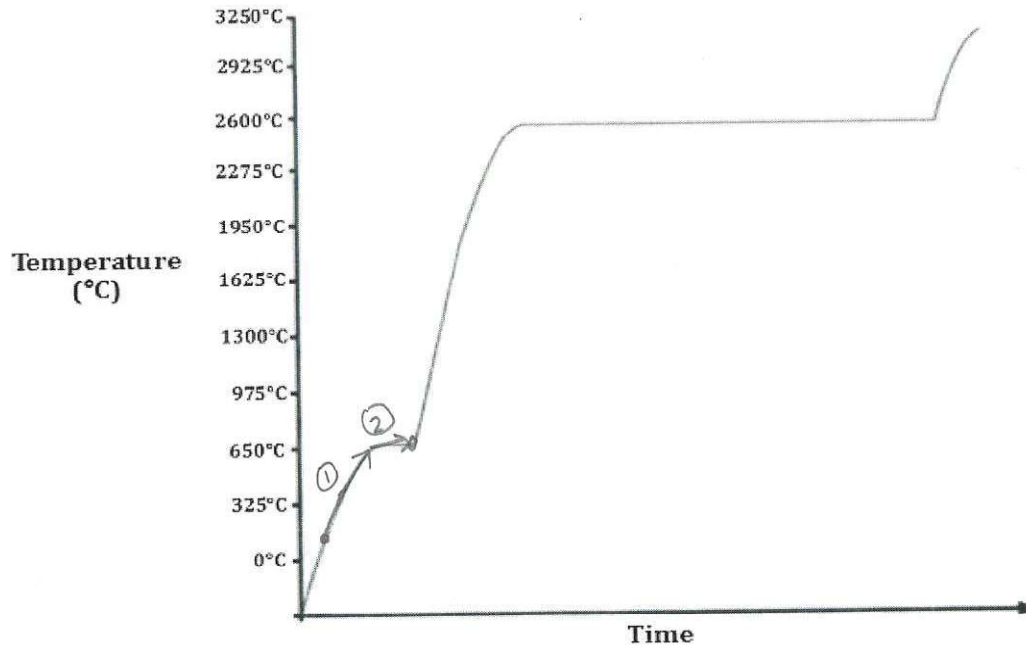
Calculate the amount of heat needed to purify 1.00 mol of Al originally at 298 K by melting it. The melting point of Al is 933 K. The molar heat capacity of Al is 24 J/mol K, and the heat of fusion of Al is 10.7 kJ/mol.

a. 15,300 kJ   b. 30,600 kJ   c. 10.7 kJ   d. 25.9 kJ

① Heat $Al_{(s)}$: 298 K → 933 K ⟹ $q = nC\Delta T = (1\,mol)(24\frac{J}{mol\cdot K})(933-298) = (24)(635)$

≈ 25 × 600

② melt $Al_{(s)} → Al_{(\ell)}$ ⟹ $q = n\Delta H_{fus} = (1\,mol)(10.7\frac{kJ}{mol})$ = 15,000 J

= 10.7 kJ | Total = 10.7 + 15 = 25.7 | = 15 kJ

9. Benzene, C_6H_6, has a melting point of 5.5°C and a boiling point of 80.1°C. Given the other constants below for benzene, which of the following processes would require the ==least== amount of energy?

$$c_{solid} = 118.4\,J\cdot mol^{-1}\cdot K^{-1}$$

$$c_{liquid} = 134.8\,J\cdot mol^{-1}\cdot K^{-1}$$

$$c_{vapor} = 82.44\,J\cdot mol^{-1}\cdot K^{-1}$$

$$\Delta H_{vap} = 30.77\,kJ\,mol^{-1}$$

$$\Delta H_{fus} = 9.9\,kJ\cdot mol^{-1}$$

a. Vaporizing 2.0 mol of benzene at 80.1°C.

b. Warming 2.0 mol of benzene from 91°C to 101°C.] gas ⟹ C = 82.44

c. Warming 2.0 mol of benzene from −10°C to 0°C.] solid ⟹ C = 118.4

d. Melting 2.0 mol of benzene at 5.5°C.